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ABSTRACT

This student guide is for the second year of the Portland Project, a three-year integrated secondary school science curriculum. "Motion and Energy," the first of the two parts in this volume, begins with the study of motion, going from the quantitative description of motion to a consideration of what causes motion and a discussion of Newton's laws. There follows the development of the laws of conservation of momentum and energy, including a discussion of energy in biological systems, and culminating with a discussion of the kinetic molecular theory. The major objective of "Chemical Reactions," part two, is to build some of those basic chemical concepts that are necessary for the analytical approach in parts three and four of the third year course: "The Chemistry of Living Matter" and "Energy Capture and Growth." Some of the topics considered in part two include the mole concept, equation writing, energetics associated with chemical reactions, the dynamic nature of particles and their interactions, and the application of energy and equilibrium to chemical systems. The Harvard Physics Project textbook and the Chemical Educational Materials Study (CHEMS) textbook are used for most of the reading assignments in the second year course. A review of the development of the three-year program, its rationale and content, and a three-year course subject outline are included. (Also see SE 013 702, SE 013 703, SE 013 705, and SE 013 706.) (Author/PR)

Year Two

MOTION AND ENERGY
CHEMICAL REACTIONS

PARTS ONE & TWO
OF
AN INTEGRATED SCIENCE SEQUENCE

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STUDENT GUIDE

1971 EDITION

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STUDENT GUIDE
MOTION & ENERGY
CHEMICAL REACTIONS

AN INTEGRATED SCIENCE SEQUENCE

1971 EDITION

DIRECTOR:

DR. KARL DITTMER
PORTLAND STATE UNIVERSITY
PORTLAND, OREGON

CO-DIRECTOR:

DR. MICHAEL FIASCA
PORTLAND STATE UNIVERSITY
PORTLAND, OREGON

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Dedication

This volume is dedicated to the memory of Vernon Cheldelin under whose guidance and leadership integration of the sciences for Oregon secondary school youth was begun in 1963.

Introduction

The Portland Project was initiated in the fall of 1962 when two secondary school teachers, one with background in CBA chemistry, the other having responsibility for PSSC physics, began to note and discuss the redundancy in their respective courses. Why should students be subjected to this repetitious and fragmented representation of the physical sciences? they asked. A Steering Committee met to pursue the problem further and perhaps enlist the support of a funding organization to permit its exploration in depth. Under the able and devoted leadership of Vernon Cheldelin, Dean of the School of Science at Oregon State University (deceased), two proposals prepared for support by the National Science Foundation were funded in the summers of 1963 and 1964.

Thirty-five scientists and teachers devoted various quantities of time as writers, consultants, pilot teachers, and evaluators, with the aim of ascertaining the feasibility and efficiency of the integration of chemistry and physics. Concurrently and subsequently, other groups in other parts of the country have carried on studies that are approximately parallel to this one.[1] Though the conceptual development and points of emphasis differ, the various groups are satisfied that integration of science courses is not only feasible but highly desirable.

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[1]Federation for Unified Science (FUSE) was recently organized to act as a clearinghouse of information on integrated science courses. Victor Showalter at Ohio University is the chairman of this committee.

Dr. Michael Fiasca of the Education and Science Staffs of Portland State University conducted an evaluation which revealed that subject matter achievement in chemistry and physics and critical thinking abilities are enhanced among students who studied the integrated courses over those who study the separate disciplines of chemistry and physics. It should be emphasized that though these differences were apparent, it could not be demonstrated that they were statistically significant.[2] A concomitant result showed that enrollments in the two-year integrated courses were dramatically greater than in the separate courses.

A survey completed April 16, 1967 showed that there were forty-four schools in twenty states using the Portland Project integrated chemistry-physics course.[3]

Mounting evidence in the literature from prominent persons working in science education strongly supported this mode of organization. Dr. Jerrold R. Zacharias, the prime instigator of the PSSC physics program, exemplified the changing attitude of scientists and educators:

> The division of science at the secondary school level, into biology, chemistry, and physics is both unreasonable and uneconomical.

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[2]Detailed results of this study may be obtained by writing to Dr. Fiasca at Portland State University.

[3]Detailed enrollment figures and addresses of people who are using the Portland Project courses may also be obtained from Dr. Fiasca.

> Ideally, a three-year course that covered all three disciplines would be far more suitable than a sequence of courses which pretends to treat them as distinct. Today such a three-year course would be difficult to fit into the educational system, but much of this difficulty might be overcome at once if such a course existed, and it might well be that present tendencies in education would soon overcome the rest.
>
> In any case, a greater coordination of the three subjects is possible even within the existing framework. It is understandable that the groups which developed the existing programs, each of which faced great problems of its own as it worked toward its goals, were reluctant to embark on the larger task of giving coherence to the sum of their efforts. With the programs now complete or approaching completion, it may be that the time has arrived for this necessary step.[4]

Stimulated by the apparent success of their original work towards this kind of integrated course, persons close to the Portland Project began to discuss extension of their work to include biology with chemistry and physics in a three-year sequence. A third proposal was prepared in 1966 and granted support by the National Science Foundation. Dr. Arthur Scott, member of the Chemistry Department at Reed College who has had deep interest in the Portland Project since its inception, graciously offered his talents, energy and time to carry on the project after Dean Cheldelin's death.

A writing conference was conducted on the Portland State

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[4]From page 52 of Innovation and Experiment in Education, a Progress Report of the panel on Educational Research and Development to the U.S. Commissioner of Education, the Director of the National Science Foundation, and the Special Assistant to the President for Science and Technology, March, 1964.

University campus during the summer of 1967 to develop materials such as teacher and student guides. Eight local pilot schools committed approximately five hundred students and twelve pilot teachers for testing and evaluation. Dr. Donald Stotler, Supervisor of Science for the Portland School District, has had an active part in this and other phases of this project.

Twenty-six persons whose functions were writing, consulting, analysis, and editing met on the Portland State campus beginning June 14, 1967, to begin preliminary work on the integrated course. Their first task was to formulate an outline that displayed logical content development utilizing concepts out of biology, chemistry and physics. Particular attention was paid to matching students' abilities, interest and maturity level with the sophistication of concepts as nearly as this was possible to do. Then the committee perused material developed by the national curriculum groups -- PSSC, Project Physics, CBA, CHEMS, BSCS and IPS -- in search of material to implement the outline they constructed previously. In the absence of appropriate materials, major and minor writing projects were initiated.

The writing committee continued its work in the summers of 1968 and 1969 with Dr. Karl Dittmer, Dean of the Division of Science, as director. Four major projects were tackled and completed: (1) extensive revisions were effected in the three-year outline, (2) the first and second year courses were revised based upon student and

teacher feedback, (3) the third-year course was developed incorporating Harvard Project Physics materials as a main vehicle, and (4) an evaluation program for the three-year course was developed.

Working Committee

Carl Bachhuber - Physics
Portland State University
Portland, Oregon

Renee Bergman - Physical Science
Roosevelt High School
Portland, Oregon

Louis Bixby - Chemistry
St. Louis Country Day School
St. Louis, Missouri

Howard Browning
Chemistry and Physics
Benson Polytechnic School
Portland, Oregon

David Cox
Chemistry and Physics
Rex Putnam High School
Milwaukie, Oregon

William Devery - Psychology
Reed College
Portland, Oregon

Steven Dimeo - Editor
University of Oregon
Eugene, Oregon

Karl Dittmer
Co-Director and Biochemistry
Portland State University
Portland, Oregon

Arleigh Dodson - Chemistry
Lewis and Clark College
Portland, Oregon

Leonard M. Dooley
Biology and Chemistry
Sunset High School
Beaverton, Oregon

Ray Ellickson - Physics
University of Oregon
Eugene, Oregon

Michael Fiasca
Co-Director, Physics
Portland State University
Portland, Oregon

Edward Fuller - Chemistry
Beloit College
Beloit, Wisconsin

Curtis D. Guthrie - Chemistry
Wilson High School
Portland, Oregon

Jean Halling - Chemistry
Beaverton High School
Beaverton, Oregon

Glen Hampshire
Chemistry and Physics
Jefferson High School
Portland, Oregon

Lois Helton - Editor
Beaverton School District #48
Beaverton, Oregon

Marshall Herron - Physics
Portland State University
Portland, Oregon

George Ingebo
Director of Research
Portland School District
Portland, Oregon

Lloyd Ingraham - Physics
Grant High School
Portland, Oregon

Henry Kilmer
Biology and Chemistry
Rex Putnam High School
Milwaukie, Oregon

Helen Koopman
Physical Science
Jefferson High School
Portland, Oregon

Edward Kormondy - Director
Comm. on Undergraduate Educ.
in the Biological Sciences
Washington, D. C.

Alfred Levinson - Chemistry
Portland State University
Portland, Oregon

David Malcolm - Biology
Portland State University
Portland, Oregon

Edmund McCollough - Chemistry
Cleveland High School
Portland, Oregon

Jack McGoldrick - Chemistry
Rex Putnam High School
Milwaukie, Oregon

Lloyd Meskimen - Physics
Roosevelt High School
Portland, Oregon

Thomas Miles - Biology
Cleveland High School
Portland, Oregon

John Myers - Biology
Portland State University
Portland, Oregon

Leon Pape - Biophysics
California State College at
Los Angeles
Los Angeles, California

Arnold Pickar - Physics
Portland State University
Portland, Oregon

Clarence Porter
General Science
Portland State University
Portland, Oregon

David Porter - Physics
Portland Community College
Portland, Oregon

Jack Sadler - Biophysics
University of Colorado Medical
Center
Denver, Colorado

Guenter Schwartz - Physics
Florida State University
Tallahassee, Florida

Arthur Scott - Chemistry
Reed College
Portland, Oregon

Joseph Sklenicka - Physics
Franklin High School
Portland, Oregon

H. Dean Smith - Chemistry
Beaverton High School
Beaverton, Oregon

Donald Stotler
Supervisor of Science
Portland School District
Portland, Oregon

Harold Wik - Science Coordinator
Beaverton School District #48
Beaverton, Oregon

John H. Wirtz - General Science
Portland State University
Portland, Oregon

Pilot Schools

The following schools have used the integrated course during all or part of the past three academic years:

Adams High School
Portland, Oregon
- Glen Hampshire
- Lloyd Meskimen
- Thomas Miles

Aloha High School
Aloha, Oregon
- Mary Lou Combs
- Elvis Dellinger
- Nelson Doeleman
- Ted Parker

Beaverton High School
- Roger Berg
- Jean Halling
- Lois Helton
- H. Dean Smith

Benson Polytechnic School
Portland, Oregon
- Howard Browning
- W. B. Chase
- W. L. Hoffman
- Michael E. Lopez

Central Catholic High School
Portland, Oregon
- Jacob A. Mosbrucker
- Peter Roerig

Cleveland High School
Portland, Oregon
- John Brown
- Edmund McCollough
- Ralph Parton

Franklin High School
Portland, Oregon
- John Neeley
- Joseph Sklenicka

Jefferson High School
Portland, Oregon
- Ronald Kawamoto
- Leslie Morehead
- Kenneth Starbuck

Parkrose High School
Parkrose, Oregon
- Donald Pearson

Rex Putnam High School
Milwaukie, Oregon
- Dennis Axness
- David Cox
- Jerry Fenton
- Henry Kilmer
- Jack McGoldrick

Roosevelt High School
Portland, Oregon
- Renee Bergman
- Kenneth Fuller
- Sue Storms

Sunset High School
Beaverton, Oregon
- Leonard M. Dooley

Wilson High School
Portland, Oregon
- Curtis D. Guthrie
- Norman Sipple

Woodland High School
Woodland, Washington
- Donald G. Fry
- George L. Stroud

Acknowledgements

Linda Anderson - Receptionist
Cecelia Benner - Secretary
Judy Bierman - Secretary
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Velvet Harris - Secretary
Ethel Hays - Secretary
Lothar Graudins - Artist
Linda Kasner - Secretary
Neomia Kendrix - Technician
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Peggy Munyer - Typist
Janice Page - Laboratory Assistant
Robert Poon - Laboratory Assistant
Dianne Poppino - Laboratory Assistant
Jean Sims - Secretary
Linda Tham - Laboratory Assistant
Jill Timmons - Secretary
Kathryn Ann Torvik - Artist
Shirley Tsang - Laboratory Assistant
Barbara Wiegele - Secretary
Carol Wildermuth - Secretary
Laura Willis - Secretary

Three-Year Course Rationale

The decision to try to develop a three-year integrated science course which would replace the traditional three courses in biology, chemistry, and physics is based on several considerations. Among them are:

(1) A conviction that modern developments have made the division of science under these three headings obsolete;

(2) A recognition that the traditional courses overlap in many areas, resulting in a great deal of duplication and repetition as in the gas laws, atomic and nuclear structure, calorimetry, and the kinetic molecular theory;

(3) A feeling that terminal students, who take no more than one year of science, deserve to get a taste of all of science rather than just one aspect, as they do in the conventional programs; and

(4) A desire to emphasize the unity in the approach to natural phenomena and the similarity in the methods, techniques, and apparatus used by scientists in all fields.

A natural question arises as to what distinguishes this course from a general science course expanded to three years. The answer is that this course does not consist of a number of unrelated topics that might be taken up in any order; rather, it treats science as a structure that proceeds from observation to the development of general principles and then to the application of those principles to more involved problems.

The emphasis in a general science course is on the results of science; the emphasis here is on the methods and reasoning by which scientists have arrived at these results.

The three-year course outline shows that a number of topics such as properties of matter, energy, heat, and certain biological concepts are discussed at the first-year level and again later in the course. This recycling is deliberate. It is intended to introduce students in a semiquantitative way to some of the significant generalizations of science and to show how these generalizations arose. These topics are treated again in the second and third years when greater facility with mathematics on the part of students makes it possible for them to understand and appreciate discussions of these topics in the succinct and precise language of mathematics.

An excessively formal and quantitative approach is avoided in the first year for several reasons. Students at this level do not extract essential meaning from such a presentation of information; furthermore, first encounters with new ideas should proceed from an intuitive, non-quantitative confrontation to one that is more quantitative. Teachers have spoken out against teaching and learning methods which substitute equations, formulas and other quantitative representations for first-hand experience, word descriptions, examples, and illustrations. These criticisms are just as valid for students who are very capable and very interested in science as they are for other students. Moreover, the mathematical

sophistication of students at this level is such that they are unable to follow most mathematical arguments as explanations for natural phenomena.

The typical science experience of most secondary school students consists of one or two years devoted to general science and biology. Few study physics and chemistry. A significant advantage to the course of study described here is that students are given a chance to study physics and chemistry at a level of rigor that is consistent with their ability and their mathematical maturity. Students who terminate their study of science at the end of one year get a significant exposure to the structure of biology, chemistry, and physics as they are presented in the latest curricular developments. Students who might not elect science beyond the first year because of lack of interest in biology may be attracted by the chemistry or physics portions of the course and elect to take an additional year or two of science. Students who are "turned on" by biology may wish to pursue further study of biochemical topics in Years II and III.

First Year Course

After considering these problems and goals, the general course outline for the first year of the course was derived. It consists of four main parts:

(1) Perception and Quantification

(2) Heat, Energy and Order

(3) Mice and Men

(4) Environmental Balance?

The year begins with a study of the perceiver, moves on to the perceived, and ends with the interaction of the perceiver with the perceived. The first-year student starts out by gaining a better awareness of the nature of his perception and senses -- the faculties that make him aware of the world around him. With an increased understanding of these perceptual abilities, he can turn to the environment and then relate himself to it. He finds that his perception is limited and that he often needs to call on technological and conceptual extensions and that even these have their limitations.

The importance of organization and classification as parts of perception is emphasized. The physical properties of matter are introduced and studied as aids in organization and classification of chemicals. The identification of unknowns by study of their physical properties and use of organized data on punch cards is the culminating experiment of the Perception unit.

Apart from the great diversity exhibited in nature, which the scientist must organize in order to comprehend, certain unifying principles are essential for deeper understanding. The most powerful of these is the energy concept, which is explored in the "Heat, Energy, and Order" unit in several of its ramifications - physical, chemical, and biological. The discussion begins by developing an experientially important energy form, viz., heat. The macroscopic aspects of heat as embodied in calorimetry are related to the microscopic in terms of

random molecular motion. This builds confidence in the idea of the atomic nature of matter, which is essential to much of the unit. Various energy conversions form the vehicle for extending and generalizing the energy concept. Nuclear energy is developed in sufficient detail to underscore its environmental and social significance. Finally, the thermodynamic limitations and implications of energy conversion are explored, ending with a view of life as a supremely artful organizer in nature, a mechanism powered by energy which creates wondrous "local order", but always at the expense of influencing its environment.

The growth of a mouse colony carries the thread of the unit "Mice and Men." As it develops, students learn many things about the concept of population. The food and water consumed and products eliminated tie the mouse colony back to the unit "Heat, Energy and Order" and point ahead to the chapter on communities and to the unit "Environmental Balance?".

The cell concept is given prime position in this unit and is used to enter topics on reproduction, embryology and maturation which are observed in the mice and other organisms. The mice selected for the original colony are such that an experiment in Mendelian genetics comes out of the observations students make as the colony develops. In most of the chapters man is an important organism and receives as much attention as the mouse although the data are often secondhand.

A rather unpleasant fact that must be faced is that as our

population increases and human activities are directed towards increasing the standard of living for this population, strains are placed upon the environment. As students discovered in "Mice and Men," the size of the community has a relation to both the quantity of the food, water, and energy required and the quantity of waste products produced. To develop the concept of a closed system and point out the necessity for environmental management, an analogy between the earth and a spaceship is made. Students are then introduced by a multi-media approach to the nature of some of our common pollutants (with emphasis upon air, water, heat, noise, and radiation) as well as their effects. Following this students are encouraged to undertake a rather detailed study of a particular type or aspect of pollution. Emphasis here is placed upon student activity, which may take any number of forms. The culminating activity centers around discussion of these special studies together with the complex relations involved within the environment. It is hoped that out of these studies students will become aware of threats which exist to man's future on this planet.

Second Year Course

The second year of the course is considerably more quantitative in its approach than the first. This is the case because (1) the students are one more year along in their mathematical preparation, (2) the students who elect to take a second year of science are more likely to exert the effort to master more difficult topics, and (3) many of the quantitative aspects of physics and chemistry are basic to an

understanding of molecular biology, which is an important part of the following year's work.

The second year consists of two parts:

(1) Motion and Energy

(2) Chemical Reactions

Year II begins with the study of motion, going from the quantitative description of motion to a consideration of what causes motion and a discussion of Newton's laws. There follows the development of the laws of conservation of momentum and energy, including a discussion of energy in biological systems. This section, which is primarily mechanics, culminates with a discussion of kinetic molecular theory.

Due to recent advances in both molecular biology and biochemistry, the descriptive approach to biology has gradually given way to one that is primarily analytical. It is now necessary, even on the high school level, for the serious biology student to have a more thorough understanding of those concepts normally embodied in the "modern" high school physics and chemistry courses. The major objective of "Chemical Reactions" is to build some of those basic chemical concepts that are necessary for an analytical study of "The Chemistry of Living Matter" and "Energy Capture and Growth."

The following subtopics of this section help in the realization of the major objective: Some of the topics discussed are the mole concept, equation writing, energetics associated with chemical reactions,

the dynamic nature of particles and their interactions and the application of energy and equilibrium to chemical systems.

Third Year Course

Year III consists of four parts:

(1) Waves and Particles

(2) The Orbital Atom

(3) Chemistry of Living Matter

(4) Energy Capture and Growth

The underlying rationale of the third year is a study of energy and its importance to life. The first thrust is to build the orbital model of the atom using, as background, waves, electromagnetism and historical models of the atom. Once the orbital model is established as a representation of the localization and directionalization of electronic energy, structural models are built to show how biopolymers are spatially arranged and experiments are done to give evidence of energy relationships. With shape, size and energy relationships of molecules established, the DNA molecule is introduced. The culmination of this work comes in the final section when photosynthesis is considered. With this topic, much that has gone before is brought to a logical focus.

These topics are most appropriately placed in the third year of the integrated sequence after students have developed some facility with basic ideas from chemistry and physics - e.g., quantitative knowledge about energy, mechanisms of chemical reactions, equilibrium, rate of reaction, the photon and wave nature of light, electrical phenomena,

and kinetic molecular theory. They should not now simply parrot biochemical processes such as photosynthesis and cell respiration but should understand the many chemical and physical principles which underlie these processes.

Time is allotted at the conclusion of Year III for individual investigation and studies.

Three-Year Course Outline

TOPIC		REFERENCE
	First Year	
Part One:	Perception and Quantification	
	I. Sensing and Perceiving	PP*
	II. Measurement, Distribution, Organization and Communication	PP
Part Two:	Heat, Energy and Order	
	I. Heat	PP
	II. Temperature and Chaos	PP
	III. Energy	PP
	IV. Nuclear Energy and Radioactivity	PP
	V. Trends in Nature	PP
Part Three:	Mice and Men	
	I. Reproduction and Development	PP
	II. Genetics	PP
	III. Genetics and Change	PP
	IV. Populations	PP
	V. Ecology	PP
Part Four:	Environmental Balance?	PP

*PP designation signifies materials produced by the Portland Project.

TOPIC		REFERENCE
	Second Year	
Part One:	Motion and Energy	
I.	Motion	HP*
II.	Newton Explains	HP
III.	Multi-Dimensional Motion	HP
IV.	Conservation	HP
V.	Energy - Work	HP
VI.	Kinetic Theory of Gases	HP
Part Two:	Chemical Reactions	
I.	The Mole as a Counting Unit	PP
II.	Combinations of Gases	PP
III.	A Useful Form of P=kDT	PP
IV.	Chemical Equations	PP
V.	Electrical Nature of Matter	CHEMS+
VI.	Basic Particles	CHEMS
VII.	Energy Effects in Chemical Reactions	CHEMS
VIII.	Rates of Reactions	CHEMS
IX.	Equilibrium	CHEMS
X.	Solubility	CHEMS
XI.	Acid-Base	CHEMS

*HP designates Harvard Project Physics material.
+CHEMS designates material derived from the Chemical Educational Materials Study.

	TOPIC	REFERENCE
XII.	Oxidation-Reduction	CHEMS
XIII.	Stoichiometry	CHEMS

Year Three

Part One: Waves and Particles

I.	Waves	HP
II.	Light	HP
III.	Electricity and Magnetic Fields	HP
IV.	Faraday and the Electrical Age	HP
V.	Electromagnetic Radiation	HP
VI.	The Chemical Basis of Atomic Theory	HP
VII.	Electrons and Quanta	HP
VIII.	The Rutherford-Bohr Model of the Atom	HP
IX.	Some Ideas from Modern Physical Theories	HP

Part Two: The Orbital Atom

I.	Atoms in Three Dimensions	PP
II.	Many-Electron Atoms	CHEMS
III.	Ionization Energy and the Periodic Table	CHEMS
IV.	Molecules in the Gas Phase	CHEMS
V.	The Bonding in Solids and Liquids	CHEMS

Part Three: The Chemistry of Living Matter

I.	Monomers and How They are Built	PP
II.	Polymers or Stringing Monomers Together	PP
III.	Polymers in 3-D or The Shape of Things to Come	PP

IV. Where the Action Is--The Active Site PP

V. How Polymers Make Polymers PP

VI. Genes, Proteins and Mutations PP

Part Four: Energy Capture and Growth

I. Energy Capture PP

II. Energy Consumption - Metabolism PP

III. Metabolism and Genes PP

MOTION AND ENERGY

A Note to Teachers

This Guide is in no way meant to supplant other Teacher's Guides. It has been produced to help bridge gaps which result from the integration. It is imperative that the participating teachers become familiar with the material of all guides.

The experienced classroom teacher knows that time is always a vital factor. One objective of this integration is time. The teacher should search for that correct emphasis which permits him to cover each subject in depth and yet not belabor it such that future materials are not reached. A "development section" and a "teacher organization sheet" are included to help the teacher find that correct emphasis. The teacher should be particularly alert to the tendency to "bog down" early in the course at the expense of materials to come later.

Symbolism has been used in the teacher's guide. The following is a list of the symbols and their meanings.

PSSC	Physical Science Study Committee
CHEMS	The Chemical Education Materials Study
PP	Portland Project
HP	Harvard Project Physics
TG	Teacher's Guide
CEx.	CHEMS Exercise
CExpt.	CHEMS Experiment
HPEx.	Harvard Project Exercise
HPExpt.	Harvard Project Experiment
R	Harvard Project Reader

A PHILOSOPHY FOR USING HARVARD PROJECT PHYSICS MATERIALS

Designed for each topic are many different activities, such as text material, demonstrations, transparencies, student activities, film loops, experiments, films, problems and reader articles. To utilize most effectively these various options a class may do one of three things: (1) select those options which are most practical or most effective for the class; (2) use nearly all options in carefully planned and prepared class meetings; or (3) let different students be responsible for different options. Each of these three methods may be used alternatively or any one might always be used or a compromise to fit the occasion and alternatives available could be arranged. Each of the three methods has drawbacks, however: (1) If only certain options are chosen the time is more flexible but the opportunities for approach to a student on a given topic through various media is narrowed. (2) If all options are available and used, time for class discussion is limited and a strain is placed on the teacher for planing and pre-class set up. (3) If each option is used by teams of students, time must be made available for reports to the rest of the class which is a problem with periods of less than 50 minutes.

We have therefore attempted to give an idea of priorities as we see them, based on activities which will carry the concept of the section. Additional activities are always helpful, and specific teacher background will help determine which activities may be chosen. As with any of the media, reading in the text may be considered as a resource, to be used when other activities don't convey information or as a supplement to the other activities. A possible priority for activities might be

laboratory experience
demonstration
student activity
film loops
film
transparency
text
home desk lab problems
reader articles
lecture

A two week time cushion has been planned in addition to time alloted for the material in the course. This cushion may be used for post lab, pre lab, or seminar discussions and testing during the fourteen week period.

Suggestions for Laboratory Procedures

A laboratory is a place where scientists look at phenomena under controlled conditions. It is a place for serious work. Always prepare for an experiment by reading the directions in the manual before you come to the lab. Make a special effort to know all precautions.

Do only the experiments approved by your teacher. If you wish to do an extension (this is encouraged), check with your teacher. This general rule is for the safety of you and your fellow students. Laboratory safety is as much an attitude as a set of rules. The lab will become a safe place for investigation if the student continually uses common sense about his safety and the safety of others. If any accident does occur, report to your teacher. What seems a minor injury may have severe consequences.

You will be asked to write laboratory reports. Opinions concerning the content of these reports vary greatly. It follows that teacher judgment will determine the type of laboratory reports you are asked to write. The following ways to improve laboratory reports are to be taken as suggestions only.

(1) Mistakes should not be erased. If there is room for the correction, the mistake should be crossed out without obliterating it and the correction made. If there is insufficient room, an extra piece of paper should be added.

(2) Spelling and punctuation are important. Sentence fragments should be avoided.

(3) The report should be carefully planned. It is best to know what type of observations should be sensed and, if possible, what regularities can be found. Planning will lead to the placement of items in a logical sequence in the report.

(4) The name of the experiment should be included.

(5) The date on which the experiment was done should be included.

(6) The names of all participants should be included and the name of the person who actually prepared the report should be designated.

(7) Some reports should include a simple statement or schematic diagram of the apparatus used in the investigation.

(8) Some reports will require a brief explanation of purpose and procedure. If these are given in the laboratory manual, they should not be included in the report. Copying items is "busy work."

(9) Nearly all experiments require taking measurements and subsequent collection of data. This must be carefully tabulated. If it is possible for you to make data tables before coming to the laboratory, you will have more time for observation, which is a major part of any laboratory experience.

(10) If computations are required to interpret results, they should be included in the report. However, if several computations of a similar nature are needed, they should be illustrated with a typical example. Mathematical equations, not arithmetical operations, should be shown.

(11) If the investigation could be altered to get better results, a statement to this effect should be included.

(12) If the investigation suggests extensions, these should be described.

(13) Reading professional reports from magazines such as The Journal of Chemical Education and Scientific American should result in better reports.

(14) Many times the most significant information about the experiment is to be found by graphing results. Whenever appropriate, graphs should be included in the report; they give a picture from which regularities can be sought. You will find the following suggestions very helpful.

(a) Always use a full sheet of graph paper.

(b) Position the ordinate and abscissa far enough from the edge of the paper to allow proper labeling.

(c) Assuming a relationship exists, the abscissa should represent the independent variable; the ordinate, the dependent variable. As an example: The distance of the gas pedal from the floorboard in an automobile would be the independent variable, plotted on the x axis, while the speed of the car would be the dependent variable, plotted on the y axis.

(d) Each axis must show units - e.g., cm/sec.

(e) Labeling of each axis should run parallel to the axis.

(f) The scale of each axis should be chosen such that the functional plot covers most of the graph paper.

(g) The name of the graph, the name of the experiment and the date of the experiment should be suitably placed on the graph.

(h) When plotting data, draw a circle around each point to indicate the uncertainty associated with the measurements.

(i) Draw the smoothest possible curve suggested by your data.

MOTION AND ENERGY

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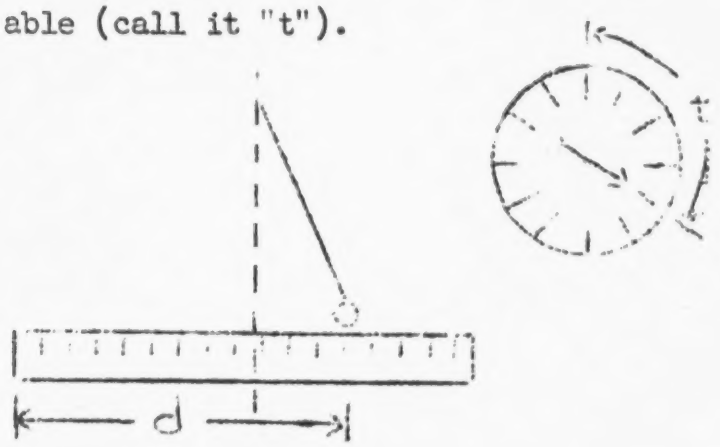
GRAPHICALLY SPEAKING . . .

A. GRAPHS AND EQUATIONS

You may recall from your math classes the meaning of the words "variable," "functional relationship," "graphical analysis," etc. The purpose of the exercises you are about to perform is to help you understand what these concepts have to do with the world of experimental science, i.e., the world outside of mathematics in which you observe, measure and predict the behavior of real objects.

First let us review some of these ideas from the viewpoint of a scientist. In some experiments we measure more than one quantity simultaneously. If there is some sort of necessary connection between the varying quantities, that is, if there is some sort of law of nature or mechanism which relates changes in one quantity to changes in another, the quantities are called <u>variables</u>. For

instance, in Year I you observed the motion of a pendulum. The position of a pendulum bob varies as time goes on. The distance of the bob measured along a ruler can be considered one variable (call it "d"), while the time observed on a stopwatch can be considered the other variable (call it "t").



The fact that the two measured quantities have some well determined connection to one another is often stated by saying that one variable is a <u>function</u> of the other, i.e., that they are <u>functionally related</u>. (In the above example <u>d</u> is a function of <u>t</u>.) There are several common ways of expressing this relationship: (1) a <u>table</u> can be made which shows values of one variable for many

corresponding values of the other; (2) a graph can be drawn in which values of one variable are plotted on one axis and values of the other are plotted on the other axis; and (3) an equation can often be found which simultaneously contains the two variables. When a functional relationship has been expressed in one of these three forms, it is often possible to transcribe this information directly into the other forms. In these exercises we will concentrate on graphs and equations. In particular we shall show how to transform the information contained (in some simple graphs such as might be obtained from experiments) into equation form.

In plotting graphs we consider one of the quantities to be the independent variable and plot it on the horizontal axis. The independent variable is usually thought of as the quantity whose value is imposed on the experiment from the outside. For instance, in measuring the performance of a racing car it might

be the amount of gas which the driver decides to feed the engine. In the case of the pendulum it would be the time t. The other quantity is thought of as representing the resultant behavior of the system and is called the dependent variable. In the case of the racing car it might be speed; in the pendulum it would be d. The dependent variable is plotted on the vertical axis. A graph of the pendulum motion would appear as shown below:

When a functional relation is expressed in equation form you all know how to obtain a corresponding graph. For practice you should try plotting the following equations: $y = 2x$; $y = 6/x$; and $y = x^2$. Consider y to be the dependent variable, and x to be the indepen-

dent variable. Note the characteristic shapes of these graphs. They are of the type you shall be most concerned with in this exercise. While the equation for the pendulum motion is a little more complicated, we shall also do that before the course is over.

The reverse process, i.e., obtaining an equation when you already have the graph, may not be familiar to you. However, it is extremely important in experimental science. Our main task now is to learn how to do this in some important simple cases.

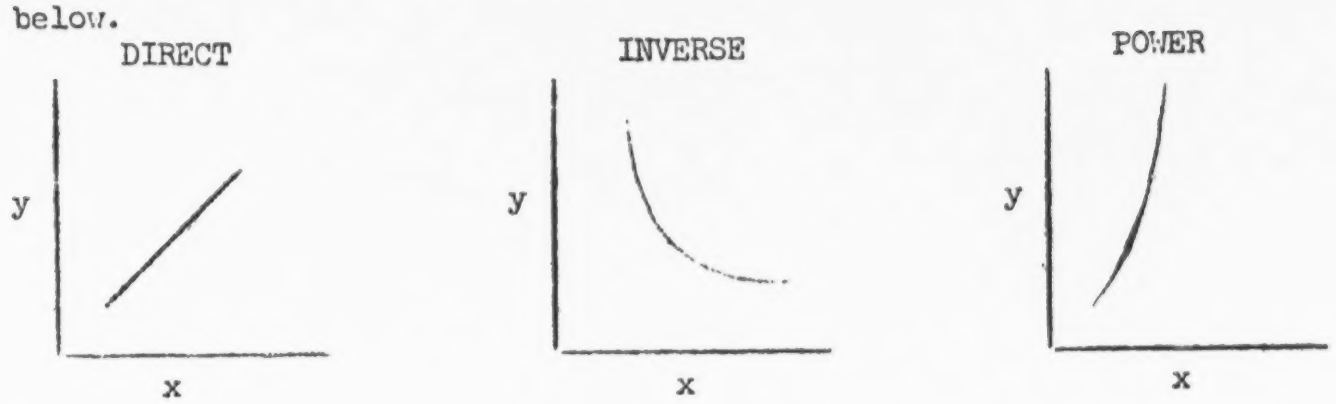
B. FROM DATA TO AN EQUATION

When experimental data is plotted directly on a graph, it may be possible to write the equation for the resulting curve. In the cases we shall be studying the process is two-fold. The first step is to discover the proportional relationship, if any, that exists between the variables. The second step is to de-

termine the value of a proportionality constant which enables us to transform the statement of proportion into a bona fide equation.

B.1 PROPORTIONAL RELATIONSHIPS

The appearance of some typical curves suggest certain proportional relationships between variables. Several common general cases are illustrated below.



In the first graph y is directly proportional to x. Symbolically this is written $y \propto x$. The plot is a straight line. (Therefore this sort of relationship is often called a "linear" relationship.)

The second graph illustrates a case where y is in inverse proportion to some

power of x. The case you are most familiar with is $y \propto 1/x$, but proportionality to $1/x^2$, $1/x^{\frac{1}{2}}$, etc. would yield curves of much the same shape. Notice how the curve bows in towards the origin and approaches the axes at the extreme ends.

The third curve is obtained in the case where y is <u>proportional to a power</u> of x. This may be written $y \propto x^n$, where the exponent n may have values such as 2, 3, etc., depending on the particular situation.* The curve approaches the origin for small values of x, but rises very steeply for large values.

B.2 THE PROPORTIONALITY CONSTANT

You probably already know that two

---

*You may have realized that the linear relationship illustrated in the first graph is actually a special case of the power relationship, where n = 1. However, this is such an important case, and yields such a distinctive curve, that it is listed as a case all by itself.

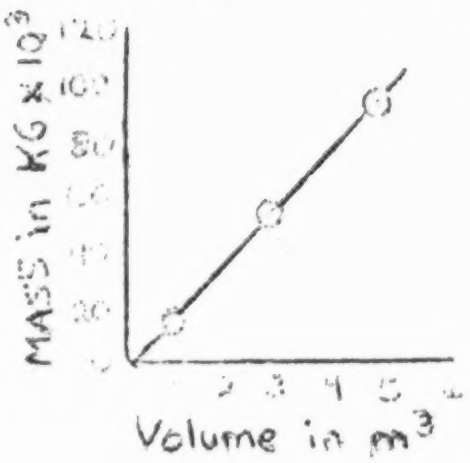
quantities which are proportional to one another may be linked together to form a correct equation by multiplying one side of the proportion by a constant of proportionality. If, for instance, $y \propto 1/x^2$, the corresponding equation could be written $y = k/x^2$. The graph would have the same shape, but it would pass through somewhat different points. When evaluated and expressed in proper units, it has special physical meaning in many experiments. In the graphs you plotted as an exercise above, the value of k for the linear relationship was 2, for the inverse relationship was 6, and for the power relationship was 1. Try plotting those equations with other values of k, say 1, 3, and $\frac{1}{2}$ respectively, in order to see what bearing changing the constant has on the curves.

C. A SIMPLE LINEAR CASE:
DENSITY OF METALS

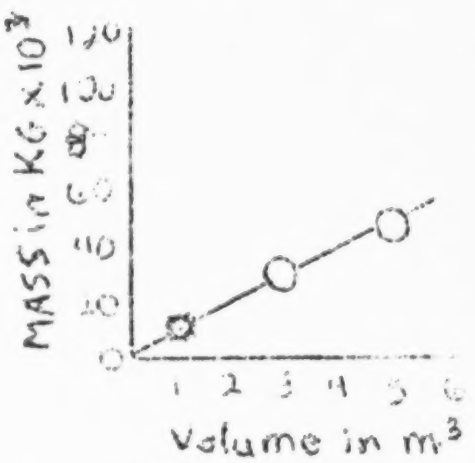
The physical interpretation of a constant of proportionality can be

easily illustrated in the case of a laboratory exercise with which you are already quite familiar. Suppose you had a number of pieces of metal, say of gold, silver, and aluminum, of various shapes and sizes. If you measured the volume and mass of each piece and plotted these quantities versus each other on graph paper, you would probably obtain straight lines similar to those shown below.

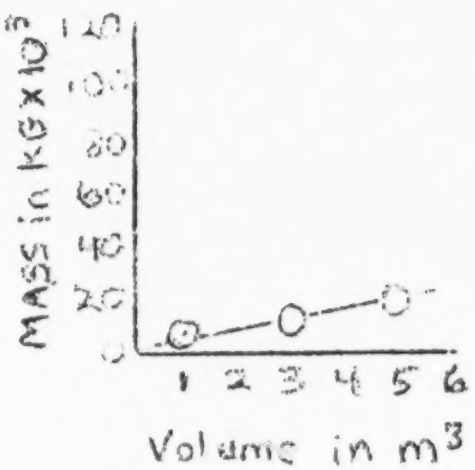
GOLD



SILVER



ALUMINUM



The basic meaning of the graphs is the same in all three cases, viz., the bigger the volume of material, the greater the mass of material, in <u>direct proportion</u>. This is of course exactly what we expect, and it is just as likely for one substance as another. But, although the

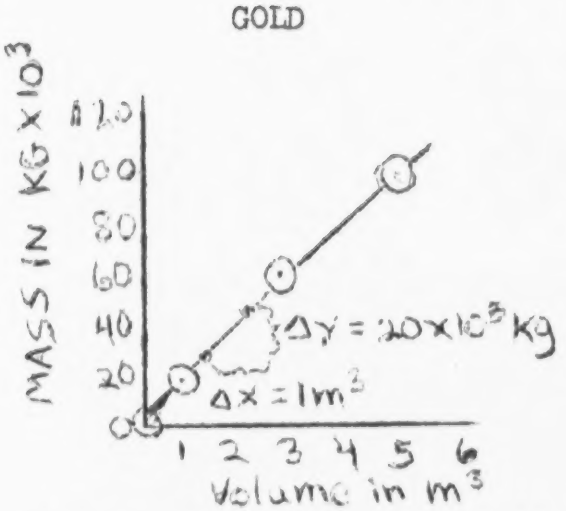
graphs are all straight lines, they differ in one important respect. The lines slant at different angles to the axis. They have different slopes. The mass of gold increases at a faster rate when the volume of material is increased than does the mass of aluminum or silver. The constant of proportionality expresses this difference for us; its value is, in fact, the slope of the graph.

The slope tells us the ratio of the change in the variable plotted on the vertical axis and the corresponding change in the variable plotted along the horizontal axis. The slope is sometimes referred to as rise/run, or designated as $\frac{\Delta y}{\Delta x}$ (where y refers to the vertical axis and x to the horizontal). Note the calculation for gold.

$$k(\text{gold}) = \text{Slope} = \frac{\Delta y}{\Delta x} = \frac{20 \times 10^3 \text{kg}}{1\text{m}^3}$$

$$= \frac{20 \times 10^3 \text{kg}}{\text{m}^3}$$

GOLD



Check yourself in this skill by finding the slope of the lines in the silver and aluminum graphs and then comparing your answers to those given below.

$k(\text{silver}) \approx 10 \times 10^3 \frac{kg}{m^3}$

$k(\text{aluminum}) \approx 3 \times 10^3 \frac{kg}{m^3}$

For the examples we have been using in this discussion, the constant of proportionality is given a special name--density. As you probably remember from Year I, density is a characteristic property of matter that tells us how the mass of any particular sample of matter is related to its volume.

In summary then, the step by step analysis of our graphical data in these

examples leads us from a statement of proportionality to a mathematical equation, and thence to an evaluation of density.

For gold:

$m \propto v$

$m = kv$

where k, the density =

$$20 \times 10^3 \frac{kg}{m^3}$$

D. ANOTHER EXAMPLE: ELECTRICAL HEATING

D.1 GETTING THE DATA

In Year I you performed experiments in which the temperature of water was raised using electrical immersion heaters. Electrical energy was put into a water bath and the consequent rise in temperature was observed. You might recall that some of the ideas you were concerned with were conservation of heat, meaning of a calorie, and conversion of energy. We will now go back and look at some data from these experiments and further analyze them in terms of graphs and equations.

In the Year I experiments styrofoam tanks were used to hold the water and the immersion heaters were operated from a D.C. power supply with provisions for varying the voltage and current. We varied the mass of the water, the length of the heating time, and the current flowing through the immersion heater, while measuring the temperature change in the water. Time was measured with an accuracy of $\pm$ 5 sec. using the sweep second hand on a wrist watch. The mass of the water was found indirectly from a volume measurement. The temperature change was measured with a Centigrade thermometer, the accuracy being $\pm$ 0.2°C. The current flowing through the heating element of the immersion heater was measured with an ammeter to within 0.01 ampere. Perhaps you would like to repeat this same experiment with equipment readily available in your laboratory. Some typical results are given below in tabular form.

Temperature Change (°C) upon Heating
Using 0.40 Amp Current

Mass(kg) / Time(sec)	0.300	0.250	0.200	0.150
60	0.4°	0.6°	0.7°	1.1°
120	1.1°	1.3°	1.7°	2.2°
240	2.1°	2.7°	3.3°	4.4°
360	3.1°	4.0°	5.0°	6.5°
480	4.2°	5.2°	6.5°	8.4°
600	5.3°	6.4°	8.0°	10.4°

Table D.1

Temperature Change (°C) of 0.200kg of
Water Heated for 300 Seconds

Current	Temperature Change
0.1 amps	0.3°C
0.2 amps	1.0°C
0.3 amps	2.5°C
0.4 amps	4.0°C

Table D.2

Notice that in the course of our experiments we were concerned with four variables (change in temperature, mass of water, heating time, and heater current),

but for the present time we will examine them only two at a time. Finally we will use all of our information to obtain a combined result.

D.2 TEMPERATURE VS. TIME

One relationship that we are interested in is that which may exist between the temperature change in the water and the length of heating time. However, we must be sure to keep the other two variables constant during this search so that our two-variable analysis will be valid. We have already learned that the first step in graphical analysis is the construction of a properly labeled graph. Plot the temperature change ($\triangle T$) against the length of heating time (t) for a mass of 0.200 kilograms of water. Remember to plot the dependent variable (temperature change in this case) on the vertical axis and the independent variable (heating time) on the horizontal axis. After plotting your

points, construct a smooth curve which fits the points as best you can estimate. Does your graph appear indicative of any of the proportional relationships that we examined earlier? If so, you may be able to state the proportional relationship that exists between the variables. As we stated earlier, to proceed from a statement of proportionality to an equation it is necessary to determine a constant of proportionality. If you have obtained a straight line, find the constant of proportionality from your graph by measuring the slope of the graph. Be sure to indicate units in your result. Now write the equation relating the two variables you have been examining

You may be curious as to whether or not this same relation applies to heating other masses of water. On the same sheet of graph paper, plot the data for each of the other masses of water. Do all the masses show the same proportion-

al relationship? Do all of the lines have the same equation? What can you conclude?

D.3 TEMPERATURE AND MASS

Now let's take a look at the relationship that may exist between the change in temperature of the water and the mass of water being heated. The time of heating and current through the heating element is to be maintained at the same value for all the measurements. Plot a graph of temperature change ($\triangle T$) versus mass of water (m) for a heating time of 120 seconds. Does your graph appear to have the same general shape as one of those already shown? Do we have a technique for finding its slope? If not, we need to try something different before we can proceed with the task of writing an equation.

Your graph probably has a shape resembling the inverse proportion curve discussed earlier. If, in fact, the re-

lationship is a simple inverse proportion $\Delta T \propto 1/m$, it is relatively easy to show this. Simply plot ΔT versus $1/m$ and a straight line should be obtained, from which a slope can be found. Try it. You should now be able to write the equation describing the relationship that exists between temperature change of the water and mass of water for a heating time of 120 seconds. You may have a suspicion that the same proportional relationship holds true for other masses of water. If you wish, go ahead and graph data for the other masses of water to find the answer for yourself.

D.4 TEMPERATURE AND CURRENT

Now that we have been reinforced by our success with the first two relationships, let's proceed on to the third. Here we are interested in the relationship that may exist between the temperature change in the water being heated (ΔT) and the quantity of current flowing

through the heating element (I). Construct your graph, plotting ΔT against I. Something new should be discovered with this graph, since it seems indicative of a power curve.

Before we can even write the correct proportional relationship between the variables, however, we need to know the value of the unknown exponent. There was actually a similar problem in the case of the inverse relation discussed in Section D.3. There we assumed the simplest possibility, viz., a simple inverse of the form $y \propto 1/x$, and came out all right. In the general case, however, there are several ways to proceed. We could look to the theory of the experiment and make an <u>educated guess</u> as to the power relation. Before the course is over, we will know enough about electricity to be able to do this, but for the present we will not go about it this way. Another approach uses slightly sophisticated mathematical techniques involving log-

arithms, but this will not be necessary here. The approach we will take, an approach which is used surprisingly often in scientific research, is that of <u>trial and error</u>. This simply means that we try plotting ΔT against different powers of I until hopefully we find a power that will give us a straight line. Plotting ΔT against I^2 is probably a good way to start. Try it. If you obtained a straight line you can now write the proportional relationship and equation. If you didn't obtain a straight line, try plotting ΔT versus I^3 or $I^{\frac{1}{2}}$, and see if that does the job. Once you have obtained a straight line, you can proceed.

D.5 A GRAND FINALE

By now you should have found the relationship between temperature change and time of heating, temperature change and mass of water, and temperature change and current in the heating element. It is

perfectly possible that in a real situation all three of the independent variables could be varying during the experiment. How then can we determine the relationship between temperature change and each of them at the same time? It's really quite easy. If the temperature change is directly proportional to the time of heating (t), the reciprocal of the mass $\left(\frac{1}{m}\right)$, and the square of the current (I^2), then it is directly proportional to the product of the three.

$$T \propto t\ I^2\left(\frac{1}{m}\right) \quad \text{or} \quad T \propto \frac{t\ I^2}{m}$$

Writing an equation from this statement of proportionality involves nothing beyond the techniques we have already utilized.

Choose any one of your three completed graphs ($\triangle T$ versus t, $\triangle T$ versus $\left(\frac{1}{m}\right)$, or $\triangle T$ versus I^2). In each case, note that two of the independent variables were held constant and one varied. Therefore you can read a temperature

change corresponding to three known values of the independent variables at any point on the graph. Using this technique, plot a graph of $\triangle T$ versus $\frac{t\ I^2}{m}$ and write the equation for the line.

E. CONCLUSION

You have now accomplished an important task. The ability to express in the form of an equation the results of laboratory research is the essence of generalization in science. This type of process has led us to many of the best known theories and principles in science, including many of those you will be studying in your first work in physics --concepts of motion.

MOTION AND ENERGY

Prologue: THE SCIENTIST AND SCIENTIFIC ACTIVITY
Volume 1 HP

I. MOTION (KINEMATICS)

A. UNIVERSALITY OF MOTION
Read sections 1.1 and 1.2 HP

B. POSITION AS A FUNCTION OF TIME
Read section 1.3 HP

C. AVERAGE SPEED
Read section 1.4 HP

D. GRAPHIC PRESENTATION
Read section 1.5 HP

D.1 Experiment: UNIFORM MOTION
(HPExpt. 4)

E. INSTANTANEOUS SPEED
Read section 1.7 HP

F. ACCELERATION BY COMPARISON
Read section 1.8 HP

F.1 LOOKING FOR LOGICAL CONSEQUENCES OF GALILEO'S HYPOTHESIS
(Read section 2.7 HP)

F.2 Experiment: NON-UNIFORM MOTION
("A 17th Century Experiment")
(HP Expt. 5)

II. NEWTON EXPLAINS MOTION

A. MASS AND FORCE
Read section 3.1 HP

B. VECTORS
Read section 3.2 HP

C. NEWTON'S LAW
Read sections 3.3 - 3.7 HP

C.1 Experiment: NEWTON'S SECOND LAW
(HP Expt.)

D. NEWTON's LAWS
Read sections 3.8 - 3.11 HP

III. MULTI-DIMENSIONAL MOTION

A. CURVED PATHS
Read sections 4.1 - 4.4 HP

A.1 Experiment: TRAJECTORIES - I

B. CIRCULAR MOTION
Read sections 4.5 - 4.7 HP

Additional Problems - Chapter 4,
Sec. 4.2 & 4.3

1. A boy standing on a bridge above the water throws a stone horizontally outward. He times the stone as taking 6 seconds to strike the water.

a) How high is the bridge?

b) How far did the stone travel horizontally before striking the water if thrown with a speed of 50 m/sec.?

2. a) What was the magnitude of displacement of the stone in the above problem after only 2 seconds?

b) What is the direction of the displacement? (Express as a ratio of vertical/horizontal distances).

3. a) What is the speed of the stone in problem 1 along the x & y axis after 2 seconds?

b) What is the magnitude of the velocity of the stone after 2 seconds?

c) What is the direction of the velocity of the stone after 2 seconds?

IV. CONSERVATION

A. CONSERVATION OF MASS
Read section 9.1 HP

B. COLLISIONS
Read section 9.2 HP

B.1 Experiment: COLLISIONS IN ONE DIMENSION
(HPExpt. 22)

C. CONSERVATION OF MOMENTUM
Read sections 9.3 - 9.7 HP

V. ENERGY - WORK

A. WORK AND KINETIC ENERGY
Read section 10.1 HP

B. POTENTIAL ENERGY
Read section 10.2 HP

C. CONSERVATION OF MECHANICAL ENERGY
Read section 10.3 HP

D. ENERGY IN BIOLOGICAL SYSTEMS
Read section 10.8 HP

E. CONSERVATION OF ENERGY
Read section 10.9 HP

VI. KINETIC THEORY OF GASES

A. KINETIC THEORY
Read sections 11.1 - 11.5 HP

A.1 Experiment: BEHAVIOR OF GASES
(HPExpt. 29)

B SECOND LAW OF THERMODYNAMICS
Read sections 11.6 - 11.8

CHEMICAL REACTIONS

I. THE MOLE AS A COUNTING UNIT

A. COUNTING THINGS

1. Exercise: COUNTING "BB'S"

B. COUNTING UNITS

C. THE MOLE

D. ATOMIC WEIGHT

1. Experiment: COPPER WIRE, SILVER NITRATE AND MOLES

2. Experiment: FROM MOLES OF SILVER TO MOLES OF SILVER CHLORIDE

E. FORMULAS FOR MOLECULES

F. MOLECULAR WEIGHTS

G. RELATIVE WEIGHTS OF MOLECULES

Chapter I: The Mole as a Counting Unit

A. COUNTING THINGS

In previous chapters a model for the gaseous state has been examined. Speeds and sizes of molecules were discussed and the ideal gas law was derived by considering topics such as pressure, momentum and temperature of gases. (P=kDT)

If man wants to communicate about particles or things, he must be able to describe them. What is the shape of a gas particle? Is it spherical? How many are there in a given volume? How does one count these extremely small particles that go dashing randomly about in a container? Let's first think about counting these gas particles. How does man usually count things and what are some alternative ways of counting things?

A.1 - Exercise: COUNTING "BB'S"

Although "BB's" are enormously large when compared to gas particles, the practice in devising an alternative

method of counting these fairly small particles will be helpful.

Count out ten BB's; don't laugh--do it. How did you count this small number of particles? Now count out one thousand BB's. Think about alternative ways of counting. What is the easiest way to count one thousand BB's? Do it. What would be the easiest way of counting all the BB's in the classroom?

Then counting particles is an easy task. All that is needed is the mass of a number of particles. But, you say, "ten gas particles mass so little (about 10^{-21} grams) that our laboratory balances cannot mass a unit of ten." What we really need is an appropriate counting unit for gas particles.

B. - COUNTING UNITS

One of the most used counting units is the dozen. We can think of the mass of twelve cows or people, but the mass of twelve atoms is very little different from the mass of ten. Obviously

the dozen is not a good counting unit to use when working with gas particles. What about the gross (144 things)? No, since a gross of gas particles masses about 10^{-20} grams. We need a "really big" counting unit.

C. - THE MOLE

Such a unit already exists. It is called the mole. It has nothing to do with little animals that live in the ground, unless you wish to count them. The mole is 6.02×10^{23} things, just as the dozen is 12 things. One can talk about a mole of cows or mosquitoes; however the use of such a large counting unit for this purpose is ridiculous. One mole of cows would mass about 10^{26} Kg! This is approximately sixteen times greater than the mass of the earth!! Gas particles, however, have such small masses that a mole of gas can be massed on a typical laboratory balance.

D. - ATOMIC WEIGHT

If our gas particle happens to be mono-atomic (composed of just one atom), then the weight in grams of one mole of the gas is its atomic weight. The weight of one mole of helium gas, for example, is four grams.

Find a table of atomic weights and determine the weight of one mole of neon, krypton, and zenon. The idea of atomic weight does not stop with just gases. The weight in grams of one mole of any mono-atomic element is the atomic weight of that element. Notice the term "weight" is used in the definition of ATOMIC WEIGHT. We know the definition should read, "the mass in grams of one mole of atoms is the atomic ~~weight~~ mass". Traditionally chemists have used the term weight when they mean mass. Although tradition can be confusing, most areas of study are bound by it; and we must simply remember the chemist usually means mass when he says "weight".

D.1 - Experiment: COPPER WIRE, SILVER NITRATE AND MOLES

In this experiment you will prepare a solution of silver nitrate. You will also weigh a length of copper wire, and after it has remained in the silver nitrate solution over night, you will weigh it again. In this and the following experiment you will count, by weighing, the number of moles of particles involved.

Copper wire is composed of copper particles. It is an element; it contains only one kind of atom. The symbol for copper is Cu. Silver nitrate is a compound which has the formula $AgNO_3$. You will find your laboratory work even more exciting if you prepare a data table before you come to lab.

Your data should include the following:

Weight of Cu wire before immersion

Weight of Cu wire after immersion

Weight change of Cu wire

Weight of vial plus $AgNO_3$

Weight of vial

Weight of $AgNO_3$

Weight of beaker plus residue

Weight of beaker

Weight of residue

Obtain a 30 cm. piece of copper wire. Clean it with steel wool. Then form a coil by wrapping the wire around a small test tube. Leave about 5 cm. for a handle.

Carefully weigh the copper wire. Also weigh a clean dry 250 ml beaker. Weigh the vial of silver nitrate provided by the instructor. Do not use tap water because of impurities. Fill the beaker about two-fifths full with de-ionized water. Add the solid nitrate ($AgNO_3$) to the water. Stir gently with a glass rod until all the silver nitrate ($AgNO_3$) crystals have dissolved. Weigh the empty vial. Silver nitrate ($AgNO_3$) solution or solid will stain your skin and clothing.

watch glass to be placed on top of beaker to keep out dust

Fig. D.1

It will be easy for your instructor to determine who is the sloppy experimenter.

Now place the copper wire in the beaker with the coil immersed in the $AgNO_3$ solution. Observe and record any changes that take place for at least five minutes. Then place the beaker, covered with a watch glass, in the designated place until the next laboratory period.

When you return to the laboratory, be very careful not to disrupt the contents of the beaker until you have had a chance to observe the contents. Record your observations.

Shake the crystals off the coil and lift the coil out of the solution.

Use a wash bottle to rinse into the beaker any crystals which tend to adhere to the coil. Now set the coil aside to dry after you have rinsed it in acetone. Weigh the coil when it is dry.

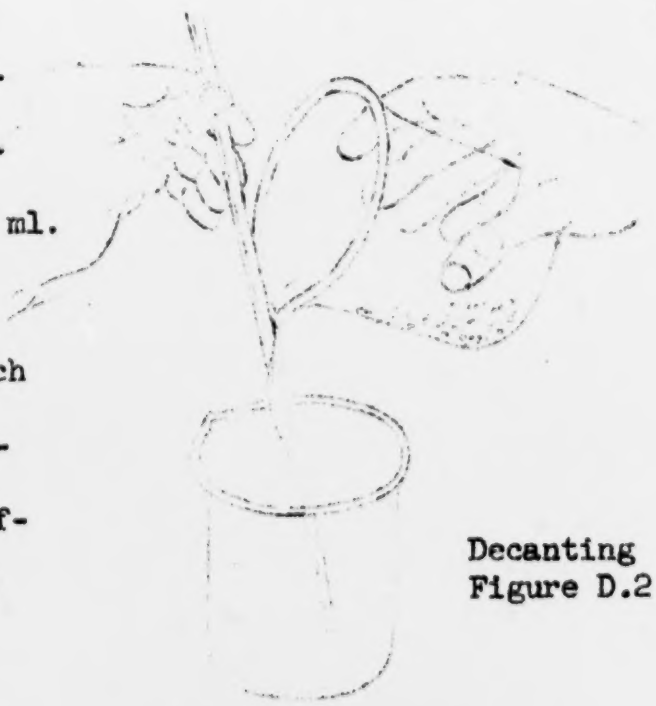
Let the crystals settle in the beaker and carefully decant the solution. Decant means to pour off the liquid and leave the solid behind.

Rinsing with a Wash Bottle

Figure D.2.A.

See Figure D.2.A. and B.

Next add 5 ml. of dilute $AgNO_3$ solution and stir gently until any Cu flecks left in the residue in the original beaker disappear. Carefully decant again. Wash the residue with 10 ml. of water and carefully decant again. You may neglect the few particles which may float over with the wash water because their weight is usually insignificant. (Uncertainty in measurement?)



Decanting
Figure D.2.B

Now the residue must be dried. Your instructor will suggest a suitable method. This usually takes less than

Drying of the sample may be carried out with a heat lamp or an oven.

24 hours.

When you return to the laboratory allow the beaker and its contents to cool before weighing. Should you use the same balance you used previously? Why? Record the weight. How can you be certain the contents are dry?

COVER THE BEAKER AND CONTENTS WITH A WATCHGLASS AND SAVE IT FOR THE NEXT EXPERIMENT.

Calculations:

1. Calculate the part of a mole of copper that went from the wire into the solution.
2. The residue is silver (Ag). Calculate the part of a mole of silver you obtained.
3. What is the ratio of moles of silver to moles of copper? Express your answer using scientific notation.
4. What you have observed can be partially described by the following: 1 mole of copper **(solid)** •

_____ mole(s) of sil-

ver nitrate (in water) ⟶

___ mole(s) of silver (solid)

+ ___ mole(s) of copper nitrate

(in water).

Using your results, write the proper whole numbers in the blanks.

5. How many atoms of Cu were lost by the wire?

6. How many atoms of silver were left as the residue?

7. What is the ratio of atoms of silver to atoms of copper (from your calculations in #5 & #6)?

8. Your instructor or you may want to do a frequency distribution for the results of this experiment. (How often did students get the same answer for the ratio of copper to silver.) If you do the frequency distribution, please consider how the range of values obtained com-

pares with the uncertainty you consider justifiable from your measurements.

What do you think causes the blue color of the solution?

D.2 - Experiment: FROM MOLES OF SILVER TO MOLES OF SILVER CHLORIDE

In your next experiment the silver produced in the previous experiment is used to form a water solution of $AgNO_3$. This is done by allowing the silver to react with nitric acid (HNO_3). A water solution of sodium chloride (NaCl) is then added to the silver nitrate solution and silver chloride is produced.

Many techniques are reviewed in this experiment and a few new ones are learned. Learn as much as you can of these techniques--careful weighing, decanting, washing, drying and filtering--because they will be used often in future experiments.

Remember to record your data carefully. Include the uncertainty and units in each measurement. You should have a data table before you come to the laboratory.

Your data should include the following:

Weight of silver from previous experiment

Weight of silver nitrate used in previous experiment

Weight of beaker #1

Weight of beaker #1 and solid $AgNO_3$

Weight of beaker #1, filter paper, and solid AgCl

Weight of beaker #2

Weight of beaker #2 and solid NaCl

Weight of filter paper

Weight of beaker #2 and solid residue

Record the weight of the beaker used in the last experiment. (The one with the solid silver in it.) Label

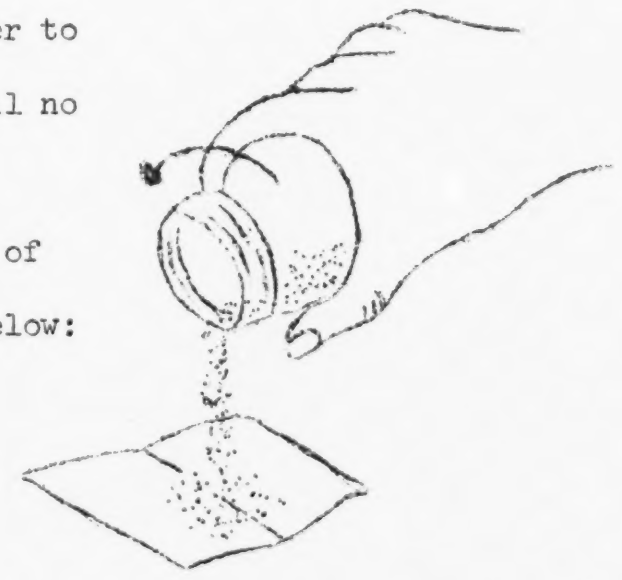
this beaker #1. Also record the weight of silver nitrate used and the weight of silver produced.

Add 15 ml. of nitric acid (HNO_3) labeled 6 M HNO_3, to the beaker containing the silver. Don't inhale any of the poisonous reddish-brown fumes of nitrogen dioxide (NO_2) which form as the silver dissolves. Put an identifying mark on your beaker (#1) and leave it in the fume hood or near an open window to dry for about 24 hours.

When you return to the laboratory, weigh the beaker which now contains $AgNO_3$. (Does this material stain clothing?)

Add 15 ml. of de-ionized water to the $AgNO_3$ in beaker #1. Stir until no more change takes place.

Pour out about $\frac{1}{2}$ teaspoonful of sodium chloride (NaCl) as shown below:



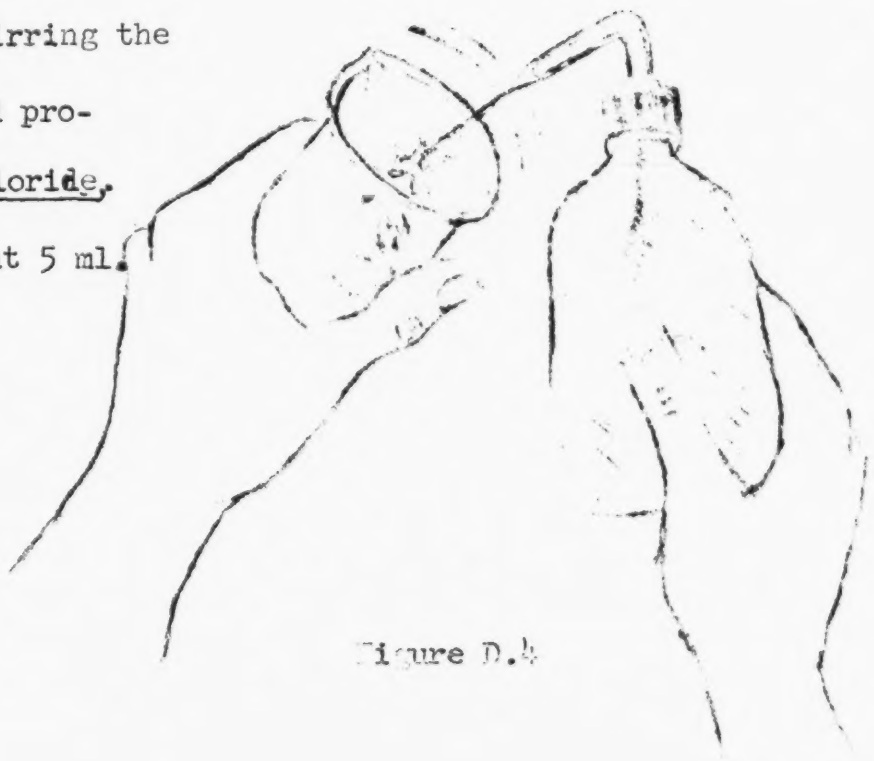
Rotate the bottle to obtain the approximate amount of solid

Now label a clean dry 100 ml. beaker as #2 and weigh it to the nearest 0.01 g.

Adjust your balance so it reads about 2.0 to 2.5 grams greater than beaker #2. Carefully add the NaCl from the paper to beaker #2 until the balance "trips." Then carefully weigh the beaker with the NaCl and record this weight on your data table.

Add about 15 ml. of de-ionized water to the NaCl in beaker #2. Stir until no more change is observed.

Add the NaCl solution slowly to the $AgNO_3$ solution while rapidly stirring the $AgNO_3$ solution. The white solid produced is the compound silver chloride, (AgCl). Rinse beaker #2 with about 5 ml. of de-ionized water.



Rinsing a beaker with a wash bottle

Figure D.4

By using two different size beakers there is less chance that the students will confuse the two solutions.

Add the rinse water to the mixture in beaker #1. Rinse beaker #2 again with de-ionized water and discard the rinse water. Beaker #2 will be used later.

Now heat beaker #1 to boiling for about 2 minutes. The solution should become reasonably clear as the precipitate settles. If you place a stirring rod in the beaker, it will help prevent unsteady boiling (bumping).

Weigh a piece of filter paper. Then fold it as shown below.

Place beaker #2 under the funnel to catch the filtrate (the liquid that comes through the filter paper). Decant the liquid from beaker #1 into the funnel. A small amount of precipitate may go into the filter paper, but try to keep most of it in beaker #1 for washing.

Wash the precipitate in beaker #1 with about 15 ml. of de-ionized water, stirring with a glass rod to aid the washing. Decant the wash water into the funnel. Repeat the washing procedure and again decant the wash water into the funnel.

After the filtration is complete place the filter paper, partially opened, in beaker #1.

Now place both beakers, #1 and #2, in a place designated by your teacher for drying. Be certain you can recognize your beakers.

When you return to the laboratory, weigh both dry samples and record the weights. (How can you be certain

the samples are dry?) Save the silver chloride as directed by your teacher.

Calculations:

Fill in the following table:

	Wt.-gms.	Moles
Ag(previous exp.)		
$AgNO_3$(previous exp.)		
$AgNO_3$ produced this exp.		
NaCl added		
AgCl (in beaker #1)		

1. How does the number of particles of $AgNO_3$ produced in this experiment compare with the number of particles of $AgNO_3$ used in the previous experiment? How do you account for any similarity or difference?
2. Compare the sum of the weights of the $AgNO_3$ and NaCl used with the sum of the weights of the AgCl and the residue in beaker #2. What is the significance

of these results? Perhaps some class data would be helpful.

3. Check on the moles of silver atoms involved in both experiments. Considering the uncertainty associated with your measurements, what can you conclude about the number of moles of silver atoms involved in this series of chemical changes?

4. Pure $AgNO_3$ is a white solid. How do you account for any color which may be present in the prepared samples?

If you have time, you may wish to devise an experiment to determine the composition of the residue in beaker #2. Check with your teacher.

E. FORMULAS FOR MOLECULES

Molecules are made of atoms in definite arrangements and definite numbers. Symbolism aids us in demonstrating the

number of kinds of atoms in a molecule. For example, H_2 means two atoms of hydrogen make up a molecule of hydrogen gas. A molecule of water is composed of two atoms of hydrogen and one atom of oxygen. The formula for water, H_2O, conveys all this information. The molecular formula for ammonia is NH_3. What does the formula represent? It is possible to find tables of the symbols commonly used for elements. The Handbook of Chemistry and Physics is a good place to look for molecular formulas.

It is often desirable to translate molecular formulas from digits to words. This is done by using the appropriate Greek prefix from the list below.

mono	one	hexa	six
di	two	hepta. . . .	seven
tri.	three	octa	eight
tetra. . . .	four	nona	nine
penta. . . .	five	deca	ten

Often no prefix is used and in these cases "mono" is understood.

Examples:

molecular formula	translation in words
CO_2	carbon dioxide
PBr_3	phosphorus tribromide
CCl_4	carbon tetrachloride
PCl_5	phosphorus pentachloride
N_2O_5	dinitrogen pentoxide
NaCl	sodium chloride

F. MOLECULAR WEIGHTS

If a mole of atoms has a definite weight, then a mole of molecules which is composed of atoms in definite numbers must also have a definite weight. One mole of H_2O must contain two moles of hydrogen atoms and one mole of oxygen atoms. Therefore the molecular weight of H_2O is 1 + 1 + 16 = 18. The gram molecular weight is the weight in grams of a mole of molecules. What is the

gram molecular weight of H_2? of NH_3?

G. RELATIVE WEIGHTS OF MOLECULES

Imagine that in your left hand you have a dozen identical red marbles that have a total weight of thirty-two grams and in your right hand a dozen identical white marbles that have a total weight of sixteen grams. The ratio of the total weight in your left hand to the total weight in your right is thirty-two to sixteen or two to one. But--can we determine the ratio of weight of one red to one white marble from this same information? Of course; the ratio of total weights is the same as the ratio of individual weights. Therefore, if the ratio of molecular weights is known, the ratio of weights of individual molecules is also known, if we know there is the same number of each. For example, one mole of oxygen weighs 32 grams and one mole of hydrogen weighs two grams. The ratio of their molecular weights

is 32 to 2 or 16 to 1. The ratio by weight of an individual oxygen molecule (O_2) to an individual hydrogen molecule (H_2) must also be 16 to 1. This comparison of the weight of a mole of one substance to the weight of a mole of a second substance yields a meaningful ratio because a mole of any substance contains the same number of molecules (6.02×10^{23}) as a mole of any other substance.

Exercises for Home, Desk and Lab (HDL)

1. A stone about the size of a grapefruit weighs roughly a kilogram. How many moles of such stones would be needed to account for the entire mass of the earth, which is about 6×10^{24} kg?
2. If we had one mole of dollars to divide among all the people in the world, how many dollars would each of the three billion inhabitants receive?
3. The gram atomic weight of gold is 197. Determine the weight in grams of one gold atom.
4. The most delicate balance can detect a change of about 10^{-8} grams. How many atoms of silver would be in a sample of that weight? The gram atomic weight of silver is 108.
5. How many moles of atoms are in
   (a) 9.0 grams of aluminum
   (b) 0.83 grams of iron?

6. What do the following symbols represent? K, Ca, Co, Pb, Sc.

7. Write formulas for:

   (a) silicon dioxide

   (b) sulfur dichloride

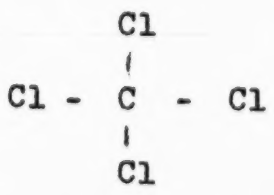
   (c) nitrogen trifluoride

   (d) aluminum tribromide

   (e) dinitrogen diiodide

8. (a) What does the molecular formula CCl_4 mean?

   (b) What information is added by the following structural formula? Guess!



9. How many moles of oxygen atoms are in one mole of perchloric acid molecules ($HClO_4$)?

10. Consider the following data:

Element	Atomic Weight
X	12.0
Y	19.0

X and Y combine to form a new substance Z. If one mole of

X combines with four moles of Y atoms to give one mole of Z, (a) what is the formula of Z in terms of X and Y? (b) what is the weight of one mole of Z?

11. A chemist weighs out 10.0 grams of carbon dioxide, 10.0 grams of water and 10.0 grams of magnesium. How many moles of each substance does he have?

12. Write the molecular formulas for the following compounds and give the weight of one mole of each: carbon monoxide, silicon dioxide, phosphorus tribromide, titanium tetrachloride, and uranium hexafluoride.

13. Calculate the molecular weight of each of the following:

(a) SiF_4

(b) HF

(c) F_2

(d) Xe

(e) NO_2

14. The ratio of the weight of a liter of chlorine gas (Cl_2) to the weight of a liter of oxygen gas (O_2) is 2.22 (both measured at room temperature and pressure). Calculate the molecular weight of chlorine.

15. Gaseous uranium hexafluoride is important in the preparation of uranium as a source of atomic energy. A flask is filled with gaseous uranium hexafluoride. The uranium hexafluoride, UF_6, is found to weigh 7.04 grams. The same flask is filled with oxygen and, under similar conditions, the oxygen weighs 0.64 grams. What is the ratio of the weight of one uranium hexafluoride molecule to the weight of one oxygen molecule?

16. Determine the weight ratio of an oxygen atom to a carbon atom.

17. Oxygen (O_2) is found to weigh 8 times as much as an unknown gas. What is the molecular weight of the unknown gas?

II. COMBINATIONS OF GASES

A. AVOGADRO'S HYPOTHESIS

A.1 Experiment: THE WEIGHTS OF EQUAL VOLUMES OF GASES AT THE SAME TEMPERATURE AND PRESSURE

A.2 THE EFFECT OF BUOYANCY ON WEIGHING

B. THE VOLUME OF A MOLE OF GAS

C. S. T. P.

D. TOTAL PRESSURE AND PARTIAL PRESSURE

D.1 Experiment: A QUANTITATIVE INVESTIGATION OF THE REACTION OF MAGNESIUM WITH HYDROCHLORIC ACID

Amadeo Avogadro, an Italian scientist who lived during the 18th and 19th century, noticed that gases combine in small whole number ratios. For example:

Two liters of the gas nitric oxide	combines with	one liter of the gas oxygen	to form	two liters of the gas nitrogen dioxide
One liter of the gas oxygen	combines with	two liters of the gas hydrogen	to form	two liters of the gas water

These combinations in small whole numbers occurred only if the volumes were measured at the same temperature and pressure. What do these observations have to do with particles and counting them?

A. Avogadro's Hypothesis

Equal volumes of gases, measured at the same temperature and pressure, contain equal numbers of molecules.

This hypothesis, first made in 1811, is still accepted today. That equal volumes contain equal numbers of particles at similar conditions allows the direct determination of relative weights of gases.

A.1 - Experiment: THE WEIGHTS OF EQUAL VOLUMES OF GASES AT THE SAME TEMPERATURE AND PRESSURE

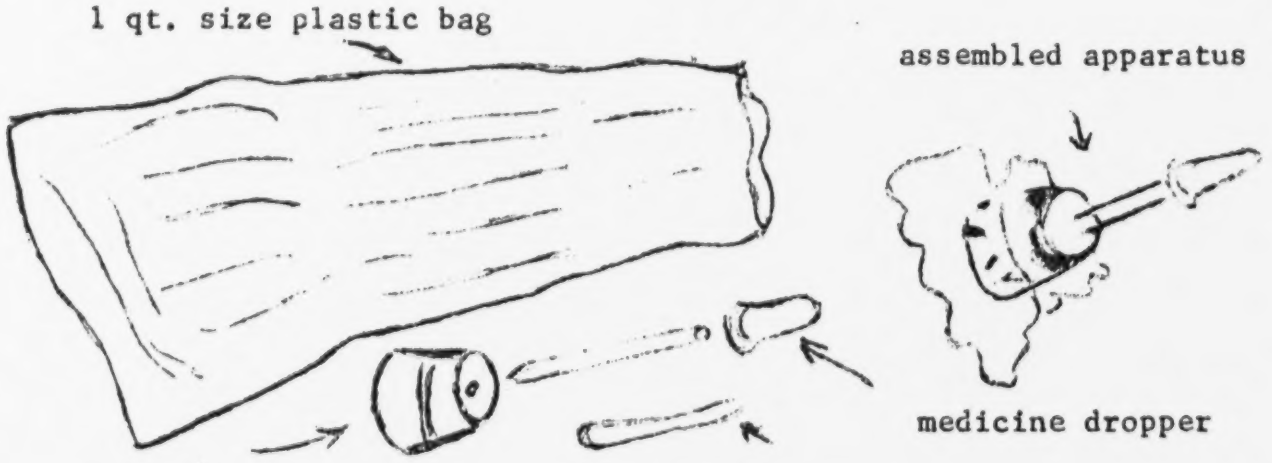
In this experiment you will weigh equal volumes of carbon dioxide and oxygen and perhaps another gas. As you know, the volume of a gas varies with temperature and pressure. Precautions must be taken to keep these variables the same as the gases are weighed. Precautions must also be taken in weighing, because the weight of the gas samples will be at most a few grams. With such small weights, the uncertainty of measurement could destroy the significance of your data. A few greasy finger prints or a moisture droplet are enough to destroy your results. In other words "cleanliness is next to good results."

Before coming to the laboratory read the experiment (pages 42 - 50) and organize a data table.

Part I. Weighing carbon dioxide gas $(CO_2)_{(g)}$

Obtain the apparatus pictured below from your teacher and put it together as pictured. Figure A. 1.

Good quality bags are required. "Glad Bags" work.



Be certain the large end of the stopper is inside the bag. If you fold the bag around the stopper in small pleats, there will be less chance of leakage. Be careful when you force the medicine dropper into the stopper. Experience shows it is less painful if there is toweling between glass and your hand if the glass breaks. Ask your teacher for some help if you do not already know the proper

Save the stopper assembly for future years.

technique for handling glass tubing.

Press out any air in the bag by smoothing it flat. Then replace the rubber cap on the medicine dropper and weigh the complete apparatus. Don't forget to record the uncertainty.

Remove the rubber cap and connect the apparatus to the $CO_{2(g)}$ source supplied by your teacher. Allow the bag to be fully inflated. Hold the apparatus by the stopper and disconnect the rubber tubing from the medicine dropper. Don't squeeze the bag; but let the excess gas escape so that the bag will be at room pressure. Now replace the rubber cap while still holding the bag.

Weigh the bag apparatus with the CO_2 gas and record, the uncertainty in the measurement. You may wish to repeat the weighing of the $CO_{2(g)}$ in order to check your work.

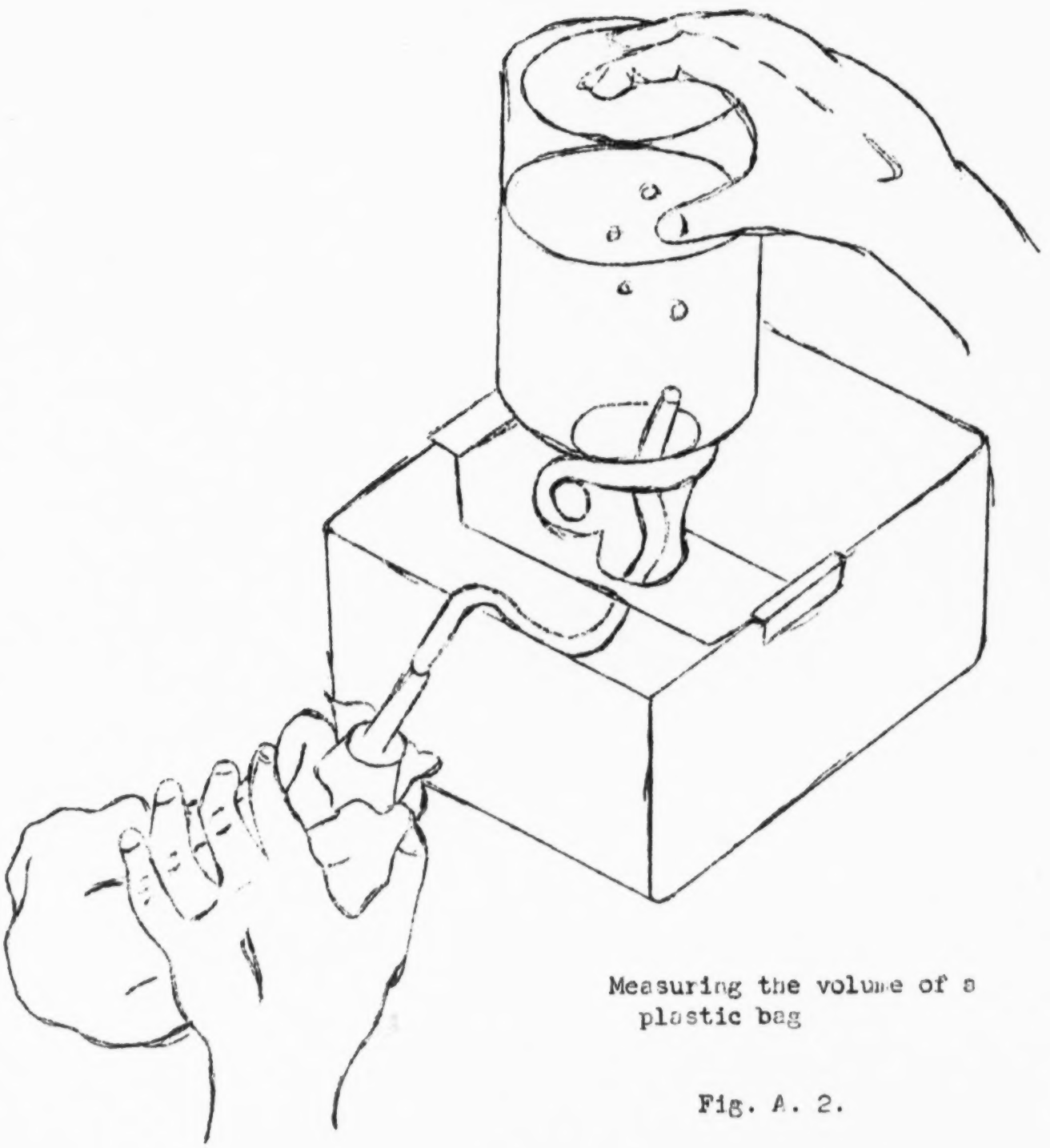
Part II. Weighing oxygen gas ($(O_2)_g$)

Empty the bag and stopper apparatus. Fill it with $O_{2(g)}$ using the source supplied by your teacher. Remember to allow the excess $O_{2(g)}$ out of the bag. Weigh the bag apparatus with the $O_{2(g)}$ and record with the uncertainty. Remember--the idea is to have equal volumes of gases at the same temperature and pressure. Save the apparatus (bag filled with oxygen) for the determination of volume that follows.

If another gas is available and you have time, you may want to obtain the weight of the third bag of gas.

Part III. Volume determination

Measure the volume of oxygen in the bag by the method shown in the picture on the following page. (Figure A.2.)



Measuring the volume of a plastic bag

Fig. A. 2.

Completely fill a large bottle or jug (about 2 liters) with tap water. Place a glass square over the mouth of the jug and invert it over the trough of water. Remove the glass slide under water.

Shove one end of a suitable piece of rubber tubing into the jug. Push it up within one centimeter of the bottom of the jug.

Now remove the cap from the medicine dropper and attach it to the other end of the rubber tubing. Gently press on the bag so the gas will displace the water in the bottle. Finally smooth out the bag to remove all the gas.

Quickly pinch the tubing so that the back pressure will not force oxygen (or water if the end of the tube is below the water level) back into the bag. Pull the tube out of the jug and set the jug upright on the table.

Measure the volume of water needed to refill the bottle. Record the vol-

ume of the bag. What is the uncertainty of the measurement? Why is the volume needed to refill the bottle the volume of the bag full of gas? Record room temperature and pressure. If you haven't read a barometer before, ask your teacher or someone who has to show you how.

A.2 - THE EFFECT OF BUOYANCY ON WEIGHINGS

Because we live in a sea of air, we usually do not consider its effects on measurements. When any object is weighed in air, it is buoyed up by a force equal to the weight of the volume of air it displaces. When you step on your bathroom scales, you read a weight that is less than your true weight. Your volume is about 70 liters. At normal conditions of temperature and pressure seventy liters of air is about 84 grams or three ounces.

(70 liters)(1.2 g/liter) = 84 grams.

Unless you have an unusual bathroom scale, the uncertainty of the instrument is much greater than three ounces. We

usually neglect the buoyancy effect because it is relatively small.

Suppose that instead of the air around us we were immersed in water and we had the proper scuba equipment. As you probably know, objects are easier to lift when they are submerged in water. The water buoys up a submerged object with a force equal to the weight of the volume of water it displaces. For example, an object with a volume of one liter (1000 ml.) when submerged in water displaces one liter of water. Therefore, it is buoyed up with a force of 1000 g., the weight of a liter of water. The same object in air is buoyed up with a force of about 1.2 grams (the weight of the liter of air it displaces).

In this experiment the apparent weight of the gas is not the true weight of the gas. We must add the weight of the volume of air displaced to the apparent weight of the gas in order to find the true weight of the gas.

The weight of the volume of air displaced can be found by using the

following table.

WEIGHT OF A LITER OF AIR IN GRAMS PER LITER, ± 0.01 g, AT VARIOUS TEMPERATURES AND PRESSURES.

Pressure (mm Hg)	Temperature 15°C	20°C	25°C	30°C
700	1.13	1.11	1.09	1.07
710	1.14	1.12	1.10	1.09
720	1.16	1.14	1.12	1.10
730	1.18	1.16	1.14	1.12
740	1.19	1.17	1.15	1.13
750	1.21	1.19	1.17	1.15
760	1.23	1.21	1.19	1.16
770	1.24	1.22	1.20	1.18

Table A.1.

Calculations

1. What is the apparent weight of oxygen that was in the bag?
2. What is the weight of air displaced by the bag?
3. What is the true weight of the oxygen that was in the bag?
4. What is the true weight of the carbon dioxide that was in the bag?

5. What is the ratio by weight of the volume of $CO_{2(g)}$ to the volume of $O_{2(g)}$?
6. What is the ratio by weight of an individual $CO_{2(g)}$ particle to an individual $O_{2(g)}$ particle?
7. Use a table of atomic weights to calculate the ratio of a mole of $CO_{2(g)}$ to a mole of $O_{2(g)}$.
8. Keeping uncertainty in mind, how do your answers to 5 and 6 compare with the answer to # 7?

B. THE VOLUME OF A MOLE OF GAS

One mole of most gases will occupy 22.4 liters if the temperature is 0° Celsius and the pressure is one atmosphere (760 mm Hg).

C. STANDARD TEMPERATURE AND PRESSURE (S. T. P.)

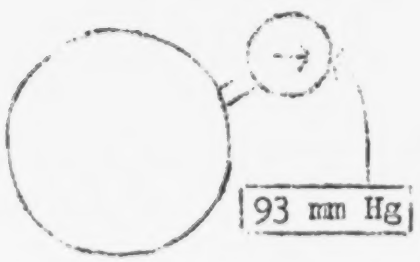
By arbitrary agreement among scientists throughout the world, standard temperature is 0° C and standard pressure is one atmosphere (760 mm Hg).

D. TOTAL PRESSURE AND PARTIAL PRESSURE

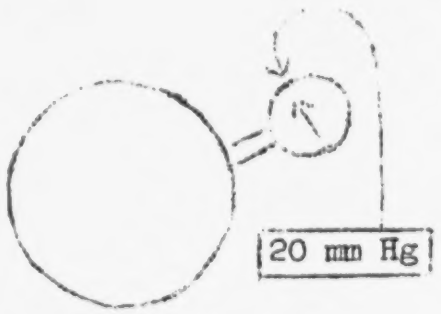
The figure below shows three bulbs of equal volume. The first bulb contains 0.0050 mole of air. The pressure gauge shows the pressure to be 93 mm Hg. The second bulb contains 0.0011 mole of water vapor. The pressure in the second bulb is 20 mm Hg. The third bulb contains 0.0050 mole of air and 0.0011 mole of water vapor. The pressure in the third bulb is 113 mm Hg. This indicates that the pressure exerted by the mixture of gases is the sum of the pressures of the separate components of the mixture.

113 mm Hg = 93 mm Hg + 20 mm Hg.

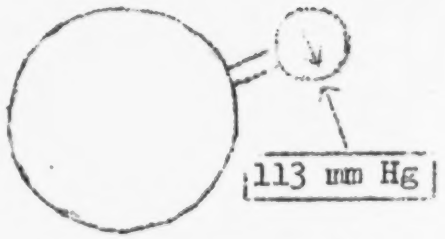
Because similar experiments yield the same results, we can state: The total pressure is the sum of the pressures exerted by each of the components of the gas mixture. The pressure exerted by each of the gases in a gas mixture is called the partial pressure of that gas. In other words, the partial pressure is the pressure each gas would exert if it were all by itself in the container.



0.0050 mole of air at standard temp.



0.0011 mole water vapor at standard temperature



0.0050 mole of air plus 0.0011 mole water vapor at standard temp.

If the 0.0050 mole of air contains 0.0040 mole of N_2 and 0.0010 mole of O_2, what is the partial pressure of O_2 in the first bulb? What is the partial pressure of O_2 in the third bulb? Don't read further until you have answered these two questions.

D.1 - Experiment: A QUANTITATIVE INVESTIGATION OF THE REACTION OF MAGNESIUM WITH HYDROCHLORIC ACID

This experiment gives you the opportunity to determine the volume of hydrogen gas produced when a small amount of magnesium reacts with hydrochloric acid. The hydrogen gas is collected over water at room temperature and pressure. The major purpose of this experiment is to determine the volume of dry hydrogen (that is, $H_{2(g)}$ without any $H_2O_{(g)}$) that would be produced by one mole of magnesium.

Obtain a piece of magnesium ribbon and carefully cut off a piece 5.00 cm long. Your teacher will give you the weight of 1.000 meter of Mg ribbon. Since the ribbon is of uniform thickness and width, you can calculate the weight of the short piece of Mg you will use. Record your data promptly (see p. 56).

Wad the Mg ribbon so that it can be encased in a cage made of fine copper wire. Leave about 5 cm of copper wire to serve as a handle.

Position a ring stand and utility clamp to hold a 50 ml gas measuring tube which has been fitted with a one-hole rubber stopper. Have a 400 ml beaker about two thirds full of tap water handy.

Now tip the gas measuring tube slightly to one side and pour in 10 ml of the hydrochloric acid which is labeled 6 M HCl. With the tube in the same position, carefully fill the tube with water from your wash bottle. While pouring, rinse any acid that may be on the sides of the tube. This will help make certain the liquid in the top of the tube contains very little acid at the very beginning. Bubbles clinging to the sides of the tube can be dislodged by gently tapping the tube.

Hold the copper coil by the handle and insert the cage with the Mg ribbon about 3 cm down the tube. Bend the remaining copper wire over the edge of the tube and clamp it there by inserting the rubber stopper. The tube should be completely filled so that the

stopper displaces some water and some water spurts out the hole.

Cover the hole in the stopper with your finger and invert the tube in the 400 ml beaker. Clamp the tube in place. The hydrochloric acid (HCl) is more dense than water and it will diffuse down and react with the metal.

After the reaction stops, wait for about 5 minutes so that the apparatus will come to room temperature. Tap the tube gently to dislodge any bubbles clinging to the side of the tube.

Now cover the hole in the stopper with your finger and transfer the tube to a large cylinder which is almost full of water <u>at room temperature</u>. The acid solution is dilute enough so that it will not hurt you if you rinse your hand with water after the transfer. Raise or lower the tube until the level of the solution inside the tube matches the water level on the outside of the tube. This permits you to measure the volume of the gases (hydrogen and water vapor) at room pressure. The pressure

inside the tube equals the pressure outside the tube (room pressure) when the levels are the same. Why?

Read and record the volume with the appropriate uncertainty.

Then remove the gas measuring tube and clean the apparatus.

Finally, record the room pressure. You may want to repeat the experiment if you have enough time.

Your data table should include the following:

Weight of 1.000 meters of Mg ribbon____

Weight of Mg sample used____

Volume of gases:_____

Temperature of water (room temperature):___

Barometer reading (room pressure):____

Partial pressure due to water vapor at room temperature·_____

The following table will help you determine the partial pressure due to water vapor at room temperature:

VAPOR PRESSURE OF WATER AT VARIOUS TEMPERATURES (To three significant figure)

Temperature (^{o}C)	Pressure (mm Hg)
20	17.5
21	18.6
22	19.8
23	21.0
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8

Calculations

1. Find the weight of the Mg you used from the grams per meter relationship and the length of your strip of Mg.
2. How many moles of Mg did you use?
3. Assuming the only two gases involved are $H_{2(g)}$ and $H_2O_{(g)}$, determine the part of the pressure due to hydrogen gas.
4. What volume (V') would the $H_{2(g)}$ occupy if it were all by itself and the pressure were 760 mm Hg? (standard pressure)
5. What volume (V'') would the $H_{2(g)}$ occupy if the pressure were 760 mm Hg, and

the temperature were 0^{o} Celsius? (standard pressure and temperature).

6. According to your findings, how many liters of $H_2(g)$ at S. T. P. would be produced by one mole of magnesium?

Exercises for Home, Desk and Lab (HDL)

1. One volume of nitrogen gas will combine with three volumes of hydrogen gas to produce two volumes of ammonia gas. Use these data and Avogadro's Hypothesis to determine how many molecules of hydrogen combine with one molecule of nitrogen. How many molecules of ammonia are produced from one molecule of nitrogen?
2. There are a number of gas mixtures that have the following combining volume ratios: one Vol. A + one B $\rightarrow$ 2 vol. C

   If a = number of atoms in a molecule of A,

   b = number of atoms in a molecule of B,

   c = number of atoms in a molecule of C,

(a) Show that if a is even, b must be even.

(b) Show that if a is odd, b must be odd.

3. One volume of hydrogen gas combines with one volume of fluorine gas to produce two volumes of hydrogen fluoride gas (as measured at the same temperature and pressure). A variety of other types of evidence suggests that hydrogen is an element and its molecules are diatomic (two atoms per molecule).

(a) Which one of the following possible molecular formulas for the substance fluorine is <u>not</u> consistent with the volumes that combine? (Use only the data given here; do not presume the molecular formula of hydrogen fluoride.)

(1) F_2 (3) H_2F_2 (5) H_2F_4
(2) F_4 (4) HF_2

(b) For each formula in part (a) that <u>is</u> consistent with the combining volumes data and the formula H_2

for hydrogen, calculate the molecular weight indicated by that formula.

(c) For each acceptable formula in part (a), predict the molecular formula for the substance hydrogen fluoride.

4. Four volumes of hydrogen fluoride gas combine with two volumes of the gas dinitrogen difluoride to form four volumes of gas X.

(a) According to Avogadro's Hypothesis how many molecules of X are produced from one moelcule of dinitrogen difluoride?

(b) If a = number of atoms in a molecule of hydrogen fluoride, b = number of atoms in a molecule of dinitrogen difluoride, c = number of atoms in a molecule of X, write the relation among a, b and c appropriate to the combining volumes given.

(c) For each of the following possible values of a and b, calculate the

required value of c.

If a is and b is then c is _____.

a	b
1	2
1	4
2	2
2	4

(d) No odd value of b is suggested in question (c). Prove that b must be an even integer.

5. Oxygen has a density of 1.429 g/L at S. T. P. and its molecular weight is 32.00. From these data, calculate the volume occupied by one mole of oxygen at S. T. P.

6. (a) What volume will a molecular weight of any gas occupy at S. T. P.?

   (b) How many particles of any gas are in a molecular weight at S. T. P.?

7. By what factor will one mole of a gas change in volume if the pressure is reduced to 740 mm Hg?

8. By what factor will one mole of a gas change in volume if the temperature is decreased to 260 degrees Kelvin?

9. Suppose that 200 ml of hydrogen is collected over water at a temperature of 26° and a pressure of 750 mm Hg. The pressure of water vapor at 26° C is 25 mm Hg. What is the pressure of the hydrogen in the dry state at 26° C?

10. Suppose that 500 ml of oxygen is collected over water at a temperature of 23° C and a pressure of 800 mm Hg. Calculate the pressure of dry oxygen.

III. A USEFUL FORM OF P=kDT

A. VOLUME, TEMPERATURE RELATIONS

B. PRESSURE, VOLUME RELATIONS

C. COMBINED GAS LAWS

Chapter III: A USEFUL FORM OF P=kDT

In Chapter Eleven of Harvard Project Physics, an ideal gas law, P=kDT was constructed. P=kDT can be further modified and hence more easily applied to problems involving volume, temperature, and pressure of gases.

The density (D) of a gas is its mass (m) divided by the volume (V) of the gas:

(1) $D = \frac{m}{V}$ where: D = density
m = mass
V = volume of gas

P=kDT can be rewritten as

(2) $P = (k)\frac{m}{V} T$

A. VOLUME, TEMPERATURE RELATIONS

Equation (2) can be rearranged to:

(3) $V = \frac{km}{P} T$

Note that the mass (m) is directly proportional to the number of particles involved, (recall Avogadro's Hypothesis) hence assuming the number of particles to be constant also assumes the mass to be constant, then V becomes directly proportional to T (the absolute temperature).

(4) $V \propto T$

The volume is directly proportional to the absolute temperature ($V \propto T$), a relationship which is referred to as Charles' Law.

The relation between gas volume and temperature can be determined experimentally. Table A. 1. on the next page shows the results of such an experiment. These results were obtained by holding the pressure and the mass of a gas constant, and varying the volume and temperature

It is not possible to obtain data at extremely low temperatures because all gases ultimately condense into liquids. However, the experimentally derived curve may be extrapolated to the point which represents zero volume. The temperature at this point is -273°C. This is designated as absolute zero. This is, theoretically, the lowest possible temperature attainable. Any lower temperature would represent a negative volume. Scientists have recorded temperatures as low as 0.0015

T vs. V for one mole of gas
(P constant at one atmosphere)

Temperature		Volume	V/T
°C	°K	liters	liters/°K
-20	253	20.9	0.0826
0	273	22.4	0.0821
100	373	30.2	0.0810

Table A.1.

When temperatures are plotted against corresponding volumes a graph such as that in Figure A.7. is obtained.

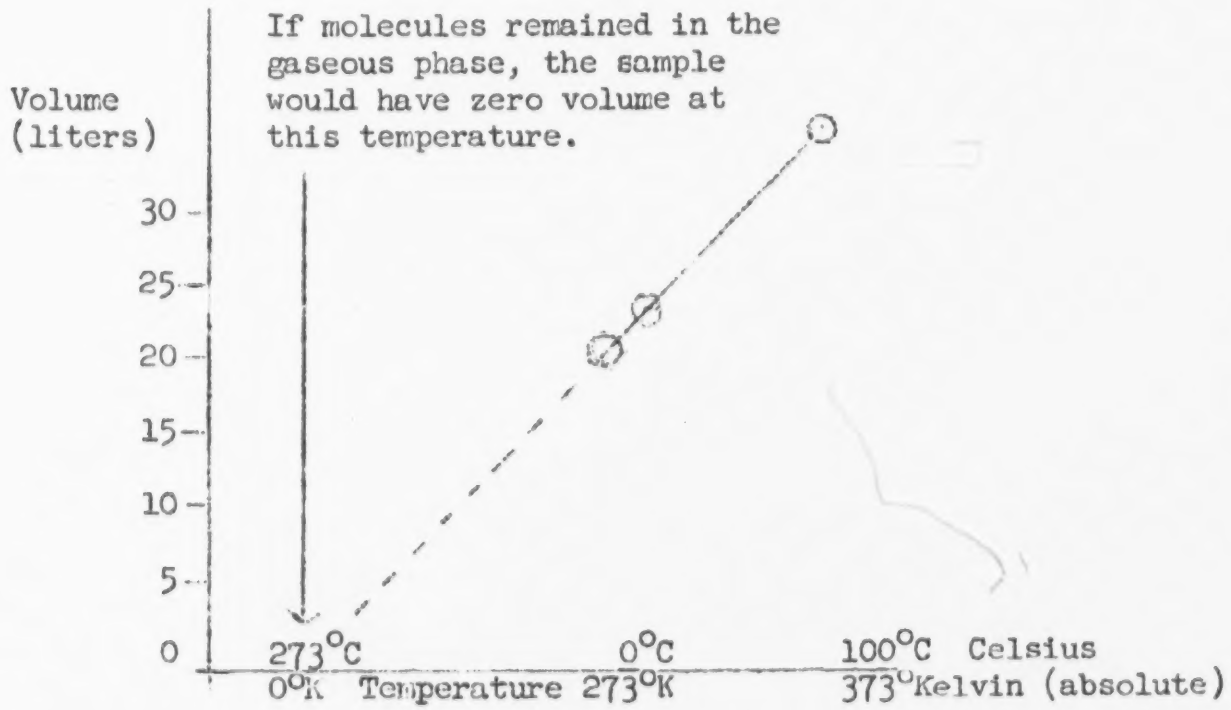


Figure A.1. - Volume of Gas vs. Temperature

degrees above absolute zero.

The Kelvin scale (°K) is a temperature scale based on an absolute zero. Look at the additional lines marked Celsius and Kelvin (absolute) below Figure A.1. How would you convert Celsius temperatures to Kelvin temperatures? You must add 273° to the °C to obtain °K. To change °K to °C you would subtract 273° from K.

T in degrees Kelvin = 273 + T in degrees Celcius.

It is important to note that the mathematical relationship which represents gas behavior is simplified by using the Kelvin temperature scale, because it is easier to work with direct proportions. Inspection of Figure A.1. shows that if we use °K for the X axis, the volume will vary by a factor of 1/273 for each degree change in temperature.

B. PRESSURE, VOLUME RELATIONS

If we hold m and T constant, then

(5) $P \propto 1/V$

The volume is inversely proportional to the pressure. ($P \propto 1/V$) is often referred to as Boyle's Law.

C. COMBINED GAS LAWS

Now let's assume we have a closed container with a movable piston (Figure A.2.) to which we will supply a definite volume of gas.

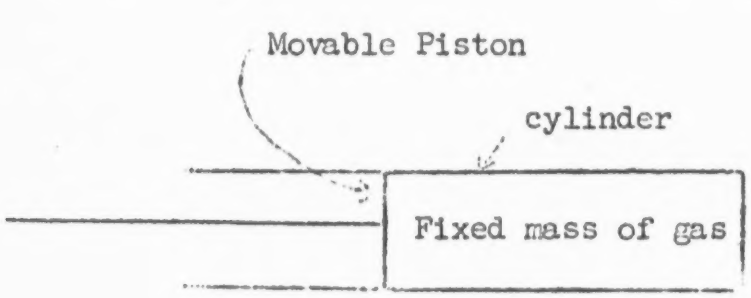


Figure A.2.

The initial conditions will be room temperature and atmospheric pressure. Our initial conditions can be designated by $\frac{PV}{T} = mk$. Now, let's move the piston (change the volume) and apply heat (change the temperature), which in each case would affect the pressure. We now have a new set of conditions for our gas sample, and we will designate these new conditions as $\frac{P'V'}{T'} = m'k'$.

When we examine the variables m and m', we observe that in both cases the mass of the gas remains constant. We did not add or subtract molecules, therefore m should be equal to m'. Since k depends on the type of gas, and we did not change the gas in the chamber, k should equal k'. It follows that

$$mk = m'k'$$

Since $mk = m'k'$:

$$(6)\quad \frac{PV}{T} = \frac{P'V'}{T'}$$

<u>This combination of both Charles' and Boyle's Laws is called the combined gas law.</u>

Let's use some algebra and rearrange our equation to find V' (the volume under new conditions). Upon rearrangement, we obtain:

$$(7)\quad V' = \left(\frac{P}{P'}\right)\left(\frac{T'}{T}\right)(V)$$

From this we can observe that we have two sets of ratios, one dealing with pressure and one dealing with temperature. Let's apply these ratios and the equation

in a sample problem.

Assuming a sample of O_2 gas in the piston apparatus occupies 100 ml at 27°C and 750 mm Hg Pressure, what volume will it occupy at standard temperature and pressure?

Solution:

1. First, itemize the given data as follows: P = 750 mm Hg

V = 100 ml

T = 27°C

but remember, it is essential to use the Kelvin temperature scale, so:

$T_K = 273 + T_C$

$T_K = 273 + 27$

$T_K = 300$

Where T_K = Temp. in degrees Kelvin

T_C = Temp. in degrees Celcius

Our problem states that the new conditions of the gas will be at standard temperature and pressure. Remember standard temperature is 0° C or 273°K and standard pressure is 760 mm Hg.

Therefore, itemize the new conditions as follows:

P' = 760 mm Hg

T' = 273°K

V' = ?

2. Use our generalizations about the behavior of gases to determine temperature and pressure ratios: If pressure increases then volume decreases; therefore, our $\frac{(P)}{(P')}$ ratio must produce a decrease of volume (V). The pressure ratio is:

$$\frac{750 \text{ mm Hg}}{760 \text{ mm Hg}}$$

If temperature decreases then volume decreases; therefore, our $\frac{(T')}{(T)}$ ratio must produce a decrease of volume. Thus the temperature ratio is:

$$\frac{273^{\circ}K}{300^{\circ}K}$$

3. To obtain the final volume (V'), multiply the initial volume (V) by the temperature and pressure fractions. The result is:

$$V' = 100 \text{ ml } \frac{(750 \text{ mm Hg})(273^{\circ} \text{ K})}{(760 \text{ mm Hg})(300^{\circ} \text{ K})}$$

$$V' = 89.6 \text{ ml}$$

4. Or we may use the combined gas law formula, $\frac{PV}{T} = \frac{P'V'}{T'}$

$$\frac{(750 \text{ mm Hg})(100 \text{ ml})}{(300^{\circ}K)} = \frac{(760 \text{ mm Hg})(V')}{(273^{\circ}K)}$$

Solving for V':

$$V' = \frac{(750 \text{ mm Hg})(100 \text{ ml})(273^{\circ}K)}{(300^{\circ}K)(760 \text{ mm Hg})}$$

$$V' = 89.6 \text{ ml}.$$

Exercises for Home, Desk and Lab (HDL)

1. (a) A thermometer containing hydrogen is at the temperature of a carbon dioxide-acetone bath (-78° C). The volume of hydrogen at standard atmospheric pressure is read. The thermometer is then moved into a bath of boiling water. The pressure on the hydrogen gas volume is still atmospheric pressure. By what factor does the volume of hydrogen change?

(b) If the thermometer had a constant volume, by what factor

would the pressure of the hydrogen gas change?

(c) If we used nitrogen gas instead of hydrogen in the thermometer, how would it affect the answers to (a) and (b)?

2. A student collects 200 ml of dry hydrogen at a temperature of 20° C. What is the volume of dry hydrogen at standard temperature if the pressure doesn't change?

3. A student collects 400 ml of oxygen at a pressure of 750 mm Hg. What is the volume of oxygen at standard pressure if the temperature doesn't change?

4. A 15.0 liter sample of dry air in a cylinder exerts a pressure of 3.00 atmospheres pressure at a temperature of 25° C. Without change in temperature, a piston is moved in the cylinder until the pressure in the cylinder is lowered to 1.00 atmospheres pressure. What

is the new volume of gas in the cylinder?

5. 100 ml of carbon dioxide gas is collected over water at a temp. of 20° C and 800 mm Hg pressure. What is the volume of dry carbon dioxide at S. T. P?

6. A 300 ml sample of nitrogen at S. T. P. is brought to 27° C and 800 mm Hg. Calculate the new volume.

7. A sample of gas occupies 275 ml at 52° C and 720 mm Hg. Correct its volume to S. T. P.

8. A sample of nitrogen is collected over water at 23° C. The vapor pressure of water at 23° is 21 mm Hg. When the pressure on the sample has been equalized against atmospheric pressure, 750 mm Hg, what is the partial pressure of nitrogen? What will be the partial pressure of nitrogen if the volume is reduced by a factor, 720/760?

9. A cylinder contains oxygen gas and a small amount of liquid water at a

temperature of 25° C. (The vapor pressure of water at 25° C is 23.8 mm Hg.) The total pressure is 700.0 mm Hg. A piston is pushed into the cylinder until the volume is halved. What is the final total pressure if the temperature of the gas doesn't change?

10. Would you expect the compressibility of solids and gases to differ? Would you expect the compressiblity of gases and liquids to differ? Why?

IV. CHEMICAL EQUATIONS

A. WRITING EQUATIONS FOR REACTIONS

B. TYPES OF CHEMICAL REACTIONS

Chapter IV: CHEMICAL EQUATIONS

A. WRITING EQUATIONS FOR REACTIONS

Scientists have a rapid and concise way of representing the relative amounts of reactants and products involved in chemical reactions. This is done by writing chemical equations. Chemical equations include the usual symbols for the elements or compounds involved and numbers representing the number of moles of each chemical.

In order to predict quantitatively the amounts of reactants needed to produce specific amounts of products, a correctly balanced equation is needed. It is obvious one must know the reactants and products if he is going to balance something involving reactants and products.

A good chemist is one who can predict the products when chemicals (reactants) are put together under given conditions. At this point in the course however, your experience is not

extensive enough for you to be able to predict the products of chemical reactions. Until you become proficient in predicting, you will be provided with the names and formulas of both reactants and products. In the meantime, you should be looking for some definite patterns as we discuss different kinds of reactions.

General "Tasks" for Writing Equations

The general tasks for writing and balancing equations are few but need to be adhered to faithfully.

Task #1 -- The first step in writing an equation for a chemical reaction is to determine the reactants used and the products formed.

Task #2 -- The second step is to write the correct chemical formula for each reactant and each product. Reactants go on the left of the equation and are separated from the products by an arrow which is read "yields."

A + B ——→ C + D

Reactant "A" plus Reactant "B" yields product "C" plus product "D".

Task #3 -- The law of conservation of mass must be satisfied. In accordance with this law, the number of atoms in the system undergoing change must be constant. A new atom cannot appear on the product side and no atom from the reactant side may disappear. You may feel the urge to change the formulas of reactants and products as you apply task #3. You must resist this urge because in changing them you are admitting you really don't have the correct formulas. You may change only the numbers that precede the formulas.

Example: Solid copper will react with an aqueous solution of silver nitrate to produce solid silver plus an aqueous solution of copper nitrate, and, if you know the respective formulas: Cu, $AgNO_3$, Ag, $Cu(NO_3)_2$, you have satisfied task #1.

Let's satisfy task #2 --

$$Cu + AgNO_3 \longrightarrow Ag + Cu(NO_3)_2$$

(Eq.1)

We now must satisfy task #3. We must be certain that our representation is consistent with our belief in the conservation of atoms. We have one mole of "Cu" on the left and one on the right. Good! Let's check the "Ag". We have one mole of "Ag" on the left and one on the right. Fine! Now let's check "N".

COPS!!!

One mole on the left but two on the right. The subscript number after the parenthesis means that everything inside the parentheses is doubled. Hence two moles of nitrogen atoms and six moles of oxygen atoms as well as one mole of copper atoms are represented by $Cu(NO_3)_2$.

Let's put "$\frac{1}{2}$" before the $Cu(NO_3)_2$. Now we have --

$$Cu + AgNO_3 \longrightarrow Ag + \tfrac{1}{2}\, Cu(NO_3)_2$$

(Eq. 2)

Now the "N" balances with one mole on each side. A quick check shows three moles of "O" on each side. But you complain, we have now represented one mole of "Cu" on the left and only $\frac{1}{2}$ mole on the right. Perhaps the equation will balance if we put "$\frac{1}{2}$" before the "Cu" on the left.

$$\tfrac{1}{2}\ Cu + AgNO_3 \longrightarrow Ag + \tfrac{1}{2}\ Cu(NO_3)_2$$

(Eq. 3)

HURRAY !!

Another check shows that we are firm believers in the conservation of mass and that we have completed all the tasks.

It should be noted here that we could have chosen 1 mole of Cu metal as the amount of reactant consumed in this reaction. We have already seen that we need twice as many molecules of $AgNO_3$ as molecules of Cu.

Therefore the reaction --

$$Cu + 2AgNO_3 \longrightarrow Cu(NO_3)_2 + 2\ Ag$$

(Eq. 4)

is also a chemical reaction in which atoms are conserved. It is just as

correct as equation (3). To show this we can multiply equation (3) by 2 in order to obtain equation (4). Remember that all coefficients in an equation may be multiplied or divided by a common factor, resulting in an equally valid equation.

There are two other tasks that will need attention as more advanced chemical reactions are studied. The first of these is that the number of electrons gained must equal the number lost in the reaction and the second is that the overall charge of the reactants must equal the overall charge of the products. These aspects of balancing equations will be considered when the moon is in the seventh house.

B. TYPES OF CHEMICAL REACTIONS

Chemists have found that most chemical reactions tend to group themselves into four general categories or types of reactions. They are as follows:

Type I: Composition (synthesis) reactions are those in which 2 or more substances react to form a more complex substance.

Composition reactions have the general form:

$$A + B \longrightarrow AB$$

Examples:

(1) Sodium + Chlorine $\longrightarrow$ Sodium Chloride

$$2Na + Cl_2 \longrightarrow 2\ NaCl$$

(2) Zinc + Oxygen $\longrightarrow$ Zinc Oxide

$$2\ Zn + O_2 \longrightarrow 2\ ZnO$$

Type II: Decomposition reactions are those in which a complex substance will break down or separate into two or more less complex substances. You will note that decomposition reactions are the reverse of composition (synthesis) reactions.

Decomposition reactions have the general form:

$$AB \longrightarrow A + B$$

Examples:

(1) Sodium Chlorate $\longrightarrow$ Sodium Chloride + Oxygen

$$2\ NaClO_3 \longrightarrow 2\ NaCl + 3\ O_2$$

(2) Mercuric Oxide ⟶ Mercury + Oxygen

$$2\ HgO \longrightarrow 2\ Hg + O_2$$

Type III: Single replacement reactions are those in which one substance is replaced in a given compound by another substance. A single replacement reaction has the general form:

$$A + BX \longrightarrow AX + B$$

Examples:

(1) Zinc + Hydrochloric acid ⟶ Zinc Chloride + Hydrogen

$$Zn + 2\ HCl \longrightarrow ZnCl_2 + H_2$$

(2) Chlorine + Sodium iodide ⟶ Sodium Chloride + Iodine

$$Cl_2 + 2\ NaI \longrightarrow 2\ NaCl + I_2$$

Type IV: Double replacement reactions are those in which two compounds interact and exchange their compenent parts to form two new compounds. A double replacement reaction has the general form:

$$AX + BY \longrightarrow AY + BX$$

Example:

(1) Sodium Nitrate + Potassium Chloride ⟶ Sodium Chloride + Potassium Nitrate

$$NaNO_3 + KCl \longrightarrow NaCl + KNO_3$$

(2) Barium Hydroxide + Hydrochloric Acid $\longrightarrow$ Barium Chloride + Water

$$Ba(OH)_2 + 2\ HCl \longrightarrow BaCl_2 + 2\ H_2O$$

Exercises for Home, Desk and Lab (HDL)

Write balanced equations for each of the following chemical reactions. The symbol "$\downarrow$" means the element or compound is a solid precipitate and the symbol "$\uparrow$" means the element or compound is a gas. A "$\triangle$" over the "yield" symbol means the addition of heat energy to the system.

(1) $Fe + O_2 \longrightarrow Fe_2O_3\downarrow$

(2) $N_2 + H_2 \longrightarrow NH_3\uparrow$

(3) $KClO_3 \xrightarrow{\triangle} KCl + O_2\uparrow$

(4) $SiO_2 + H_2F_2 \longrightarrow SiF_4\downarrow + H_2O$

(5) $Ag + HNO_3 \longrightarrow AgNO_3 + NO\uparrow + H_2O$

Balance the equations for each of the following reactions. Begin on the basis of one mole of the substance underscored.

(6) $Na + \underline{I_2} \longrightarrow NaI$

(7) $C_2H_2 + O_2 \longrightarrow H_2O\uparrow + CO_2\uparrow$

(8) $H_2 + \underline{O_2} \longrightarrow H_2O\uparrow$

(9) $\underline{Ca} + Cl_2 \longrightarrow CaCl_2$

(10) $NO_2 + H_2O \longrightarrow \underline{HNO_3} + NO\uparrow$

Write a balanced equation and indicate the reaction type (composition, decomposition, single or double replacement) for each of the following word equations:

(You may wish to consult the Handbook of Physics and Chemistry in order to obtain the correct formulas for the reactants and products involved in these reactions.)

(11) Sodium + Water ⟶ Sodium + Hydrogen Hydroxide

(12) Calcium hydroxide + Phosphoric acid ⟶ Calcium Phosphate + Water

(13) Barium Nitrate + Sodium Sulfate ⟶ Barium Sulfate + Sodium Nitrate

(14) Ammonium Sulfate + Potassium Hydroxide ⟶ Potassium + Water + Ammonia Gas Sulfate

(15) Chlorine + Magnesium Bromide ⟶ Magnesium Chloride + Bromine

Propane, a commercially bottled gas, has the formula C_3H_8. When it is burned completely the reaction products are carbon dioxide and water vapor.

(16) Write the balanced equation for the combustion of propane.

(17) One mole of propane vapor produces how many moles of water vapor when burned?

(18) How many moles of oxygen would be needed in order to burn 0.2 moles of propane?

(19) How many moles of propane would be needed to produce one mole of carbon dioxide?

(20) How many molecules of oxygen would be needed in order to burn one mole of propane?

V. ELECTRICAL NATURE OF MATTER

A. PURE SUBSTANCES
Read sections 5-1 through 5-1.3 CHEMS

B. SOLUTIONS
Read sections 5-2 through 5-2.7 CHEMS

C. ELECTRICAL EFFECTS
Read sections 5-3.1 through
5-3.5 CHEMS

D. ELECTRICAL PROPERTIES OF CONDENSED PHASES
Read sections 5-4.4 CHEMS

D.1. Experiment: REACTIONS BETWEEN IONS IN AQUESOUS SOLUTION
(CExpt. 11)

VI. BASIC PARTICLES

A. STRUCTURE OF THE ATOM
Read sections 6-1 through 6-1.5
CHEMS

B. THE SIMPLEST CHEMICAL FAMILY -
THE INERT GASES
Read sections 6-2 through 6-2.4
CHEMS

C. THE ALKALIES
Read sections 6-3 through 6-3.2
CHEMS

D. THE HALOGENS
Read sections 6-4 through 6-4.3
CHEMS

E. HYDROGEN - A FAMILY BY ITSELF
Read sections 6-5 through 6-5.2
CHEMS

F. THE THIRD ROW ELEMENTS
Read sections 6-6 through 6-6.2
CHEMS

G. THE PERIODIC TABLE
Read sections 6-7
CHEMS

VII. ENERGY EFFECTS IN CHEMICAL REACTIONS

A. HEAT AND CHEMICAL REACTIONS
Read sections 7-1 through 7-1.4
CHEMS

A.1 Experiment: A STUDY OF REACTIONS
(CExpt. 12)

A.2 Experiment: THE HEAT OF REACTION
(CExpt. 13)

B. THE LAW OF CONSERVATION OF ENERGY
Read sections 7-2 through 7-2.4
CHEMS

C. THE ENERGY STORED IN A MOLECULE
Read sections 7-3 through 7-3.2
CHEMS

D. THE ENERGY STORED IN A NUCLEUS
Read sections 7-4 through 7-4.1
CHEMS

VIII. RATES OF REACTIONS

A. FACTORS AFFECTING REACTION RATES
Read sections 8-1 through 8-1.5
CHEMS

A.1 Experiment: A STUDY OF REACTION RATES
(CExpt. 14)

B. THE ROLE OF ENERGY IN REACTION RATES
Read sections 8-2 through 8-2.4
CHEMS

IX. EQUILIBRIUM

A. QUALITATIVE ASPECTS OF EQUILIBRIUM
Read sections 9-1 through 9-1.7
CHEMS

A.1 Experiment: CHEMICAL EQUILIBRIUM
(CExpt. 15)

B. QUANTITATIVE ASPECTS OF EQUILIBRIUM
Read sections 9-2 through 9-2.4
CHEMS

X. SOLUBILITY

A. SOLUBILITY: A CASE OF EQUILIBRIUM
Read sections 10-1 through 10-1.4
CHEMS

B. AQUEOUS SOLUTIONS
Read sections 10-2 through 10-2.6
CHEMS

B.1 Experiment: DETERMINATION OF THE SOLUBILITY PRODUCT CONSTANT OF SILVER ACETATE
(CExpt. 16)

XI. ACID - BASE

A. ELECTROLYTES - STRONG OR WEAK
Read sections 11-1 through 11-1.3
CHEMS

A.1 Experiment: THE HEAT OF SOME ACID-BASE REACTIONS
(CExpt. 17)

B. EXPERIMENTAL INTRODUCTION TO ACIDS AND BASES
Read sections 11-2 through 11-2.7
CHEMS

B.1 Experiment: DETERMINATION OF THE HYDROGEN ION CONCENTRATION OF SOLUTIONS USING INDICATORS
(CExpt. 18)

C. STRENGTHS OF ACIDS
Read sections 11-3 through 11-3.5
CHEMS

C.1 Experiment: REVERSIBLE CHEMICAL REACTIONS
(CExpt. 19)

XII. OXIDATION - REDUCTION

A. ELECTROCHEMICAL CELLS
Read sections 12-1 through 12-1.4
CHEMS

A.1 Experiment: INTRODUCTION TO OXIDATION REACTIONS
(CExpt. 20)

A.2 Experiment: ELECTROCHEMICAL CELLS
(CExpt. 21)

B. ELECTRON TRANSFER AND PREDICTING REACTIONS
Read sections 12-2 through 12-2.5
CHEMS

C. BALANCING OXIDATION-REDUCTION REACTIONS
Read sections 12-3 through 12-3.3
CHEMS

D. ELECTROLYSIS
Read section 12-4
CHEMS

XIII. STOICHIOMETRY

A. A PATTERN FOR STOICHIOMETRIC CALCULATIONS
Read section 13-1
CHEMS

B. THE MANUFACTURE OF SULFURIC ACID
Read sections 13-2 through 13-2.5
CHEMS

B.1 Experiment: REACTIONS BETWEEN IONS
(CExpt. 22, Parts I and II)

B.2 Experiment; QUANTITATIVE TITRATION
(CExpt. 23)